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ERRATA

Volume 40, Number 1, February, 1947

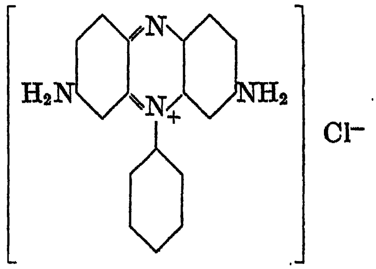
Page 146, lines 12–14: The two sentences "Thus, products marketed as 'sodium tetraphosphate' are mixtures. This has been demonstrated in other ways (10, 22)." should be transposed to the end of the following paragraph.

Page 175: In line 5 and in the last column heading of table 9, substitute "partially hydrated" for "commercial".

Communication from the author: Since publication of this review, one phosphate-manufacturing firm has informed the author that its sodium "tetraphosphate" is the only such commercial product available in the United States, that it is made by rapidly chilling from a temperature above that of fusion to prevent attainment of equilibrium at ordinary temperatures, and that the product is a clear, transparent glass. Two samples from this manufacturer were examined by the author and found to be mixtures containing some crystalline material embedded in considerable glass.

Volume 40, Number 2, April, 1947

Page 281: The formula for safranine should be changed to:



Page 313: Substitute NH_2 for NH in the formula on the right-hand side of the first equation.

Page 331: In the fourth line of the first paragraph in Section D,2 substitute "γ" for "β" before the word "picolines".

ERRATUM

Volume 39, Number 2, October, 1946

Page 256: The author wishes to give credit for figure 7B to Dr. E. O. Wiig.

ERRATUM

Volume 40, Number 1, February, 1947

Page 63: The figure given below should be substituted for the one appearing on page 63:

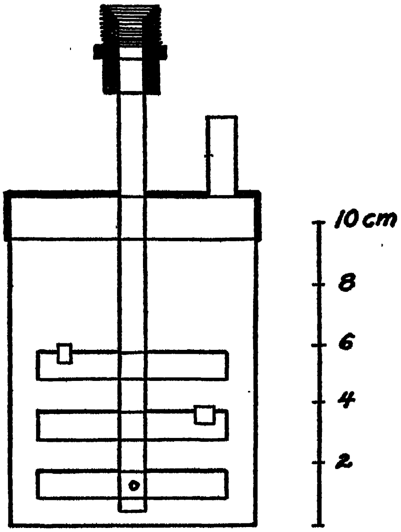


FIG. 1. Apparatus for the direct fluorination of organic compounds in carbon tetrachloride solution (Bockemüller)

ERRATUM

Volume 40, Number 1, February, 1947

Page 75: The figure given below should be substituted for the one appearing on page 75:

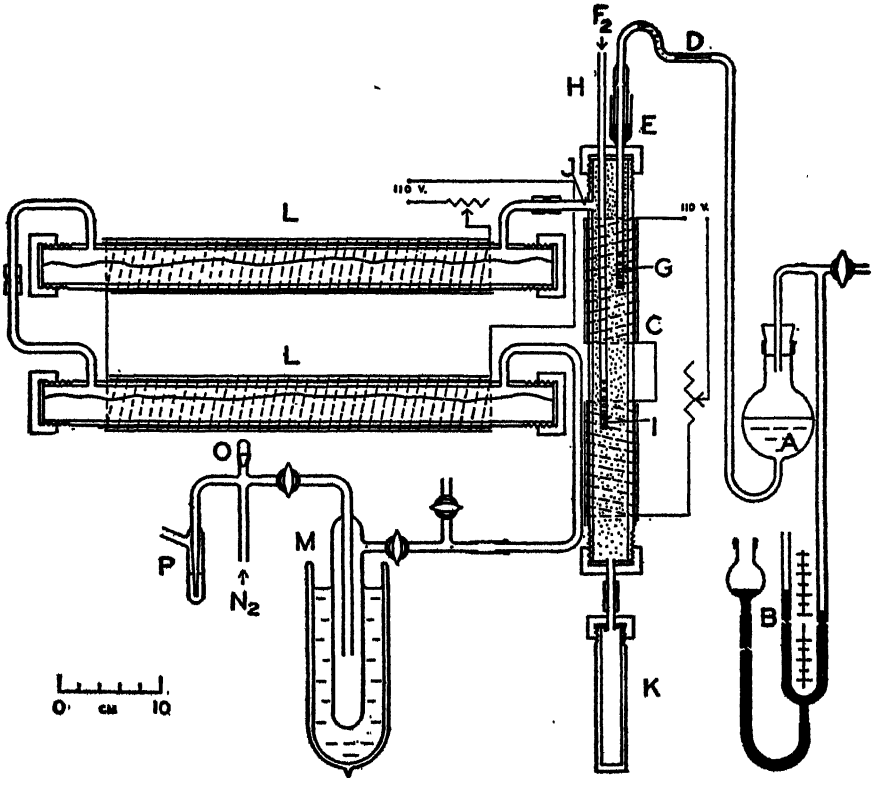


FIG. 2. Apparatus for the vapor-phase fluorination of volatile organic liquids.

ERRATUM

Volume 40, Number 1, February, 1947

Page 111: The figure given below should be substituted for the one appearin on page 111:

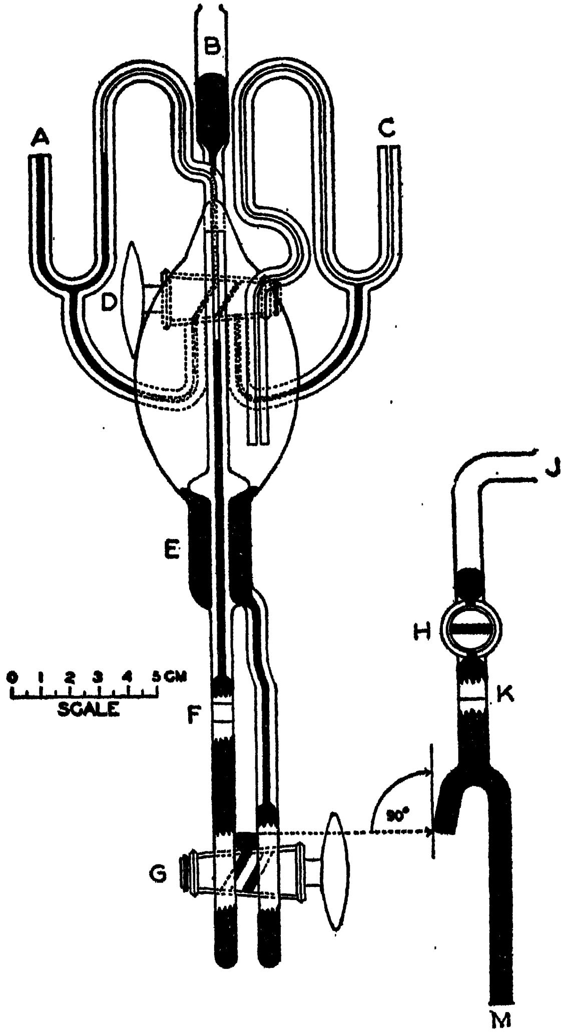


Fig. 3. Constant-volume gas buret for the quantitative estimation of elementar fluorine.

LOW-TEMPERATURE CALORIMETRY AT THE BARTLESVILLE STATION OF THE BUREAU OF MINES[1, 2]

HUGH M. HUFFMAN

Petroleum Experiment Station, Bureau of Mines, Bartlesville, Oklahoma

Received October 30, 1946

In the present review an attempt has been made to survey briefly the utility of thermodynamic data and to illustrate the methods in use at the Bartlesville Laboratory of the Bureau of Mines to obtain entropy data.

INTRODUCTION

Thermodynamic data are useful and important both for the solution of practical problems in chemistry and as an adjunct for theoretical studies of molecules. The general utility of thermodynamic data may be illustrated by the following relations of thermodynamic functions. The free energy, ΔF, is related to the equilibrium constant of a chemical reaction by the expression

$$\Delta F = -RT \ln K \tag{1}$$

In the general reaction

$$a\mathrm{A} + b\mathrm{B} = c\mathrm{C} + d\mathrm{D} \tag{2}$$

the equilibrium constant is given by the following expression

$$K = \frac{\mathrm{C}^c \times \mathrm{D}^d}{\mathrm{A}^a \times \mathrm{B}^b} \tag{3}$$

where A, B, C, and D represent the thermodynamic concentrations of the reactants at equilibrium and the products, and a, b, c, and d are the number of moles of reactants and products taking part in the reaction. By combining equations 1 and 3 the following equation is obtained:

$$\Delta F = -RT \ln \frac{\mathrm{C}^c \times \mathrm{D}^d}{\mathrm{A}^a \times \mathrm{B}^b}$$

If it is assumed as a first approximation that the thermodynamic concentration is equivalent to the actual concentration (or pressure), it is readily seen that this equation serves as a useful tool for calculating the yield of a given reaction. If the free energy ΔF is known as a function of temperature, the relative concentrations at any given temperature and pressure may be calculated. This fact is of considerable importance to the chemist, as it allows him to determine, with-

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[1] Presented at the Symposium on Low-Temperature Research, which was held under the auspices of the Division of Physical and Inorganic Chemistry at the 109th Meeting of the American Chemical Society, Atlantic City, New Jersey, April, 1946.

[2] Published by permission of the Director of the Bureau of Mines, U. S. Department of the Interior.

out resort to experiment, whether a postulated reaction is practical and also the optimum conditions of temperature and pressure for maximum yields.

The free energy is also related to certain other thermodynamic quantities by the expression:

$$\Delta F = \Delta H - T\Delta S \tag{4}$$

The interesting thing about this expression is that it permits determination of ΔF from purely thermal quantities. If studies of certain properties are made over a wide temperature range, ΔF may be evaluated at any desired temperature in this range.

The Bureau of Mines has long been interested in obtaining data that would be generally useful to the petroleum and other industries. For this reason a laboratory was established in 1943 at the Petroleum Experiment Station of the Bureau of Mines in Bartlesville, Oklahoma, for the purpose of determining the thermal data required for the solution of equation 4 over a wide temperature range.

This report is concerned with only a single phase of the general program: namely, the evaluation of ΔS, the change in entropy from low-temperature calorimetric data in conjunction with the Third Law of Thermodynamics. The Third Law may be stated simply as "The entropy of all perfect crystals at absolute zero is zero." It can be shown that the entropy of a substance which is a perfect crystal at absolute zero is given by the expression

$$S = \int_0^{T'} \frac{C_p}{T}\,dT + \frac{\Delta H'}{T'} + \int_{T'}^{T''} \frac{C_p}{T}\,dT + \frac{\Delta H''}{T''} + \int_{T''}^{T'''} \frac{C_p}{T}\,dT \cdots$$

where C_p is the heat capacity at constant pressure and ΔH is the change in enthalpy accompanying an isothermal change in state, i.e., transition, fusion, vaporization. If this expression is evaluated from 0°K., an absolute value of the entropy is obtained.

In practice, heat-capacity measurements are made from sufficiently low temperatures and an extrapolation method is employed to cover the gap between 0°K. and the temperature at which measurements are started. In the Bureau, measurements in low-temperature work are started at a temperature of about 12°K., which is attained by converting liquid hydrogen to the solid by reducing the pressure by means of large vacuum pumps.

HYDROGEN LIQUEFIER

Liquid hydrogen is not commercially available, so it was necessary for the Bureau to set up a plant for its production. Several types of hydrogen liquefiers are in use in this country, all of which involve precooling of the high-pressure gas with liquid air or nitrogen, followed by regenerative cooling to the temperature at which a portion of the hydrogen is liquefied by Joule–Thomson expansion. After consideration of various types, it was decided to adopt in principle the modified Kapitza liquefier described in detail by Blanchard and Bittner (1).

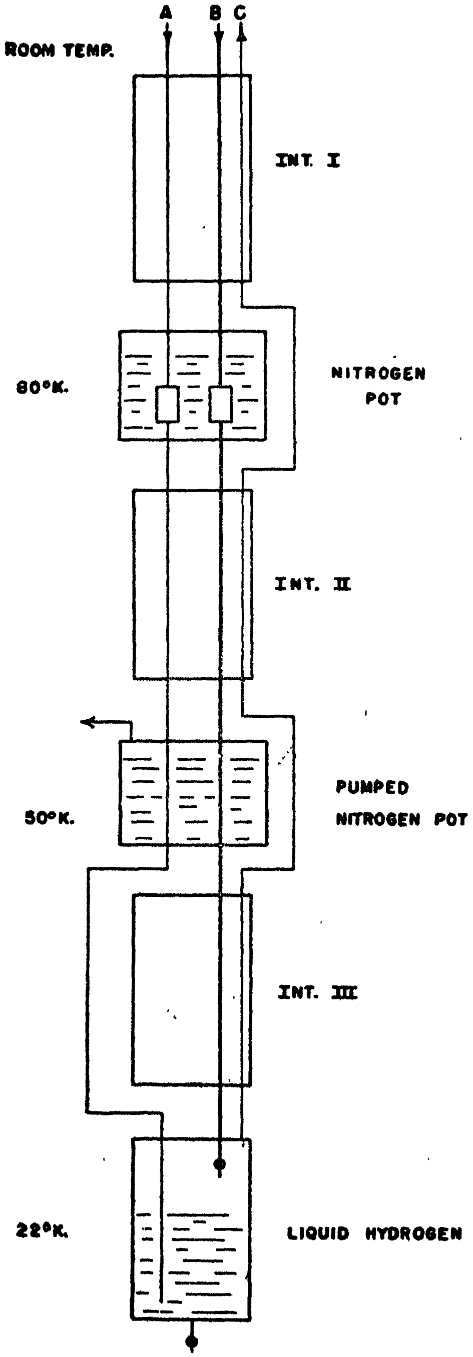

Fig. 1. Schematic flow diagram of the Kapitza scheme for hydrogen liquefaction. feed hydrogen; B, compressed hydrogen; C, exhaust hydrogen.

The principal advantage of this system is that relatively impure commercial hydrogen may be used for liquefaction.

The Kapitza principle involves two cycles: (*1*) a high-pressure cycle in which pure hydrogen is liquefied by Joule–Thomson expansion and (*2*) a low-pressure cycle in which commercial feed hydrogen is liquefied by condensation at the expense of the liquid produced in the high-pressure system. Essentially, the system is simply a method for purifying the feed hydrogen so that it may pass through the high-pressure system without plugging.

Since the system has been described in detail by Blanchard and Bittner (1), only a brief description will be given here. Figure 1 shows schematically the flow system of the liquefier. High-pressure hydrogen from the compressor enters the system at B, passes through interchanger I to coils in the liquid-air pot, where it is cooled to the temperature of liquid air boiling at atmospheric pressure, and then over activated charcoal for the removal of nitrogen. This hydrogen then passes through interchanger II to coils in the pumped air pot, where it is cooled to approximately 55°K., and then through the regenerative interchanger to the expansion valve.

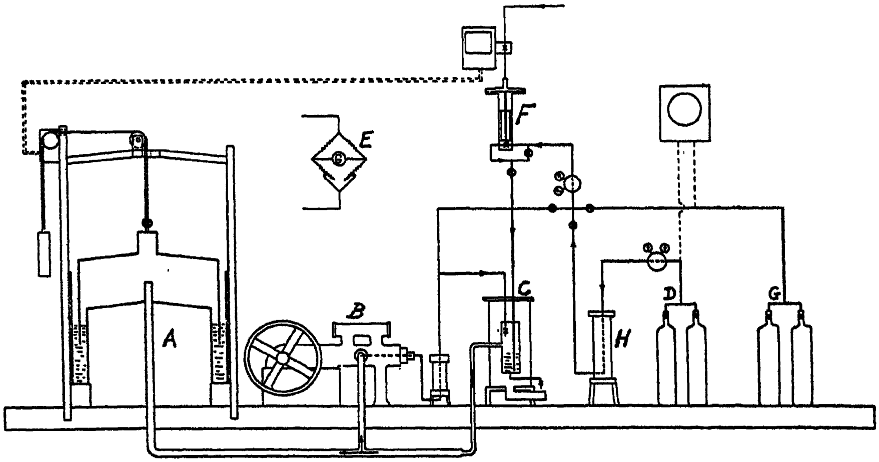

FIG. 2. Simplified flow chart of hydrogen-liquefying system

The commercial feed gas is passed over a heated copper catalyst to remove oxygen and over activated alumina to remove water; it then enters the interchanger system at A (figure 1) and passes in parallel over the same path as the high-pressure gas, except that it goes directly from the pumped pot to the liquid-hydrogen receiver. Here it is passed through the accumulated liquid and the impurities are frozen out. The purified exhaust gas from the receiver is interchanged with the incoming gas up to room temperature and then enters the gasometer on its way to the high-pressure system.

In figure 2 is shown a simplified flow chart of the liquefying system, in which only the principal units are indicated. The compressor (B) draws pure hydrogen from the gasometer (A) and delivers it to the liquefier (C). When a sufficient quantity has been liquefied, the flow of feed hydrogen stored in the cylinders (D) is started. As more hydrogen is liquefied the level of the gasometer will tend to drop. The gasometer bell is connected to the counterbalance by a chain. The chain passes over a gear, which drives a slide wire mounted in one arm of a wheatstone bridge, shown at E. The change in resistance is picked up by a micromax controller, which operates a pneumatic valve (F) in the feed line. By this arrangement the feed gas is automatically supplied at the same rate as it is liquefied. The purified hydrogen is stored in cylinders (G) when the system is not in operation. The furnace containing the copper catalyst for removing oxygen from the feed gas is shown at H.

It would be very gratifying if it were possible to say that this system was trouble free. Actually, there has been trouble from time to time with the formation of plugs in both the high-pressure system and the low-pressure system. It is believed that one of the sources of the trouble is the location of the charcoal tubes. They are so mounted in the interchanger system that it is impossible to reactivate them properly without dismounting the liquefier.

Figure 3 is a detailed cross-section drawing of the liquefier and interchanger system. This is shown only to give an idea of the complexity of the system and indicate the care that must be given to details of construction. For example, there are many silver-soldered joints, all of which must be vacuum-tight since the system is thermally isolated from the environment by evacuating the entire space within the two brass cans which contain the system.

Figure 4 is a photograph showing a general view of the liquefier and accessories. In the foreground is shown the three-stage Norwalk compressor. The control panel and the assembled liquefier are in the background. Figure 5 is a close-up view of the assembled liquefier and shows clearly the Dewar tube through which liquid hydrogen is transferred to the glass Dewars in which it is carried to the cryostats.

The main control panel is shown in figure 6. At the left of the panel is shown the micromax recording controller, which automatically controls the supply of feed hydrogen to the liquefier. The next instrument is a recording pressure gauge with two pens. One pen records the pressure in the cylinders containing the feed hydrogen. As liquefaction proceeds the pressure drops, and a record is made which is related to the amount of hydrogen liquefied. The second pen gives a record of the supply of purified hydrogen. The third instrument is a recording rotamatic flowmeter and gives a record of the flow of the gas returning from the liquefier to the gasometer or compressor. The instrument on the right is a controller which maintains the catalyst furnace, for removing oxygen from the feed hydrogen, at a constant temperature.

Below the instruments are shown the various valves and gauges which are required to control the flow and to enable the operator to know what is going on in the liquefying system. For example, there are three gauges in the high-pressure system, located at strategic points, that allow the operator to locate plugs when they are formed. The complete flow system is quite complicated and has many connections so that the gas may be directed in any chosen way. An example of the necessity for this is that gas from the compressor may be sent either to the liquefier or to high-pressure storage.

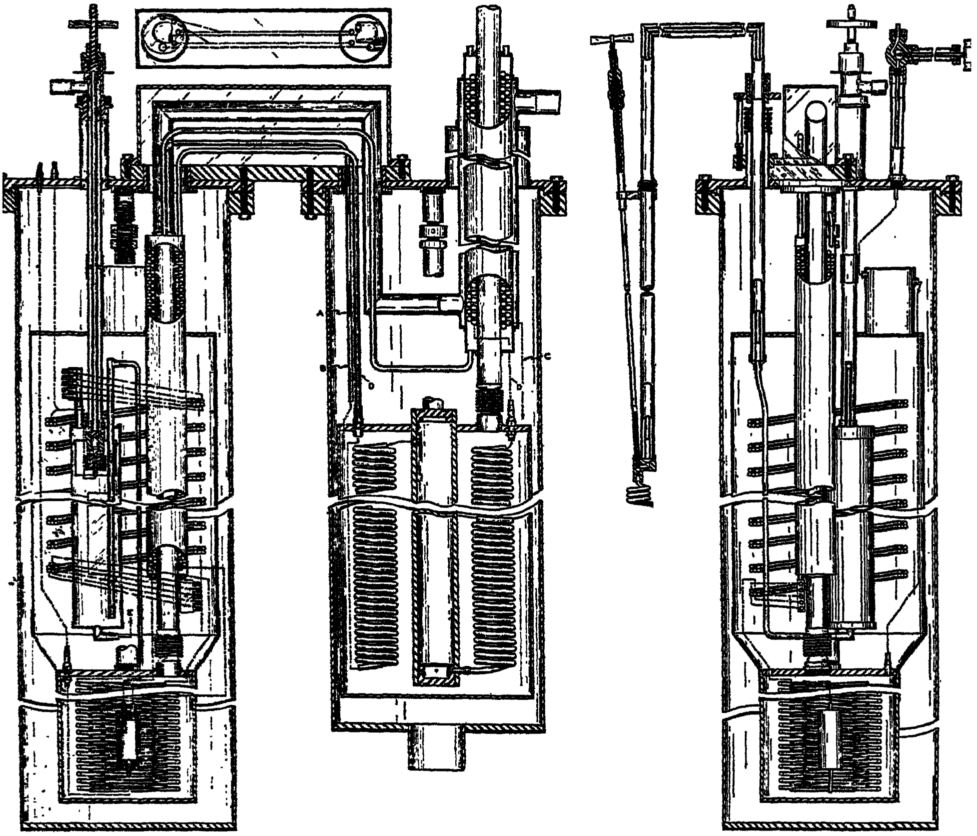

FIG. 3. Detailed cross section of liquefier and interchanger assembly. A, feed hydrogen; B, low-pressure feed hydrogen from cooling coils in nitrogen pot (these coils not shown); C, low-pressure feed hydrogen to cooling coils in nitrogen pot (these coils not shown); D, 2500 lb. per square inch; E, to liquid-hydrogen transfer tube.

The vacuum cans in which the liquefier and interchangers are mounted are equipped with blow-out plugs for safety in case the high-pressure system should open up in this closed space. There is also a safety valve in the low-pressure return line to take care of any surges that may occur across the throttling valve.

In addition to the hydrogen liquefier discussed above, the station also has its own plant for liquefying air. This is a commercial installation having a capacity of approximately 15 liters per hour.

FIG. 4. General view of room in which hydrogen is liquefied

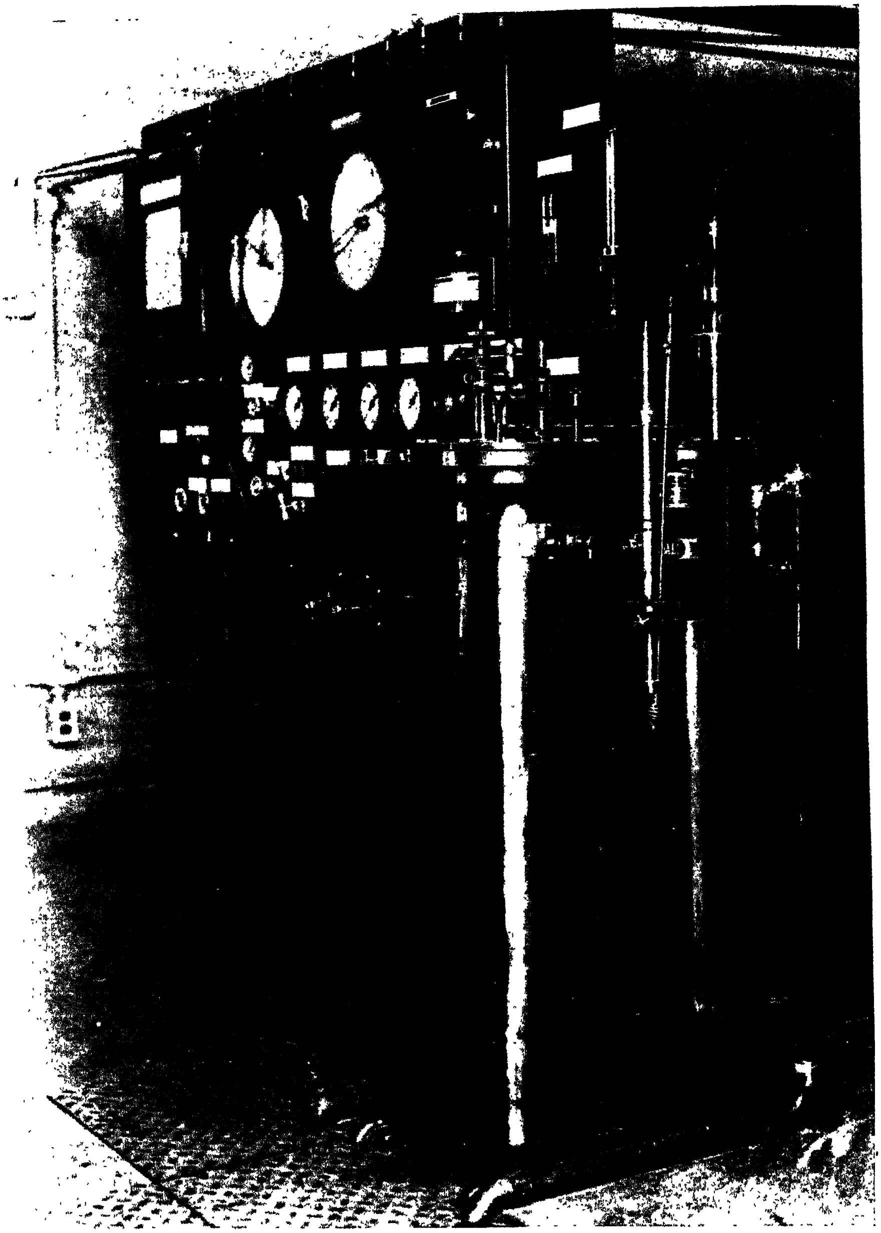

FIG. 5. Close-up view of liquefier–interchanger assembly

THE CRYOSTAT AND CALORIMETER

The cryostat is similar to that described by Blue and Hicks (2), except that a simpler device is employed to transfer heat from the calorimeter during the cooling operation. Figure 7 is a schematic drawing of the apparatus, and figure 8 is a photograph of the unassembled cryostat, showing the various tanks, radiation shields, etc. In this figure j and i are chromium-plated copper reservoirs for liquid air and liquid hydrogen, respectively, and d, e, and f are chromium-plated, thin-walled, copper radiation shields. The reservoirs are provided with thin-walled Monel-metal tubes (l, m, and n), which serve for filling, emptying, and pumping as well as for supports. The whole assembly fits into a removable brass can (g) which is connected to the top at k by a joint which can be made vacuum-tight by means of a rubber gasket and screws. This joint is always at room temperature and, consequently, is easy to make and maintain tight. The space enclosed by the can is evacuated by means of an oil diffusion pump backed by a Cenco Hyvac pump.

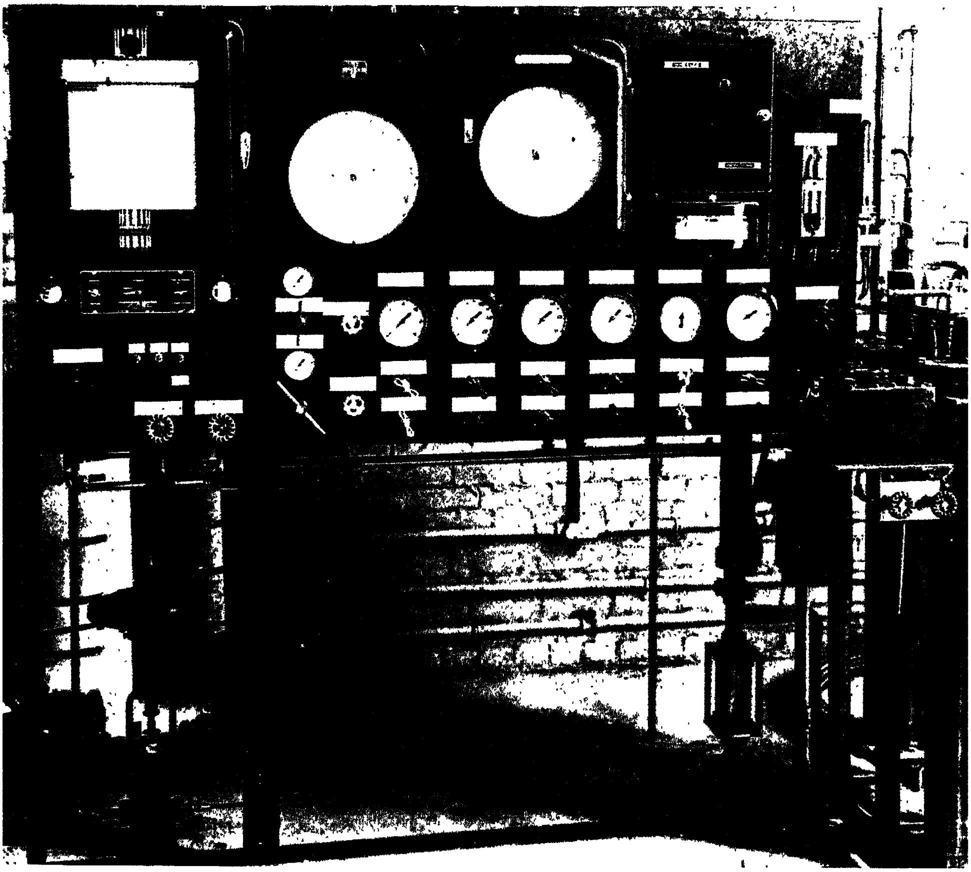

FIG. 6. Main control panel showing controls

The calorimeter is surrounded by a lightweight, chromium-plated, copper adiabatic radiation shield consisting of three parts,—top, tube, and bottom. Each part of the adiabatic shield is provided with a closely wound non-inductive

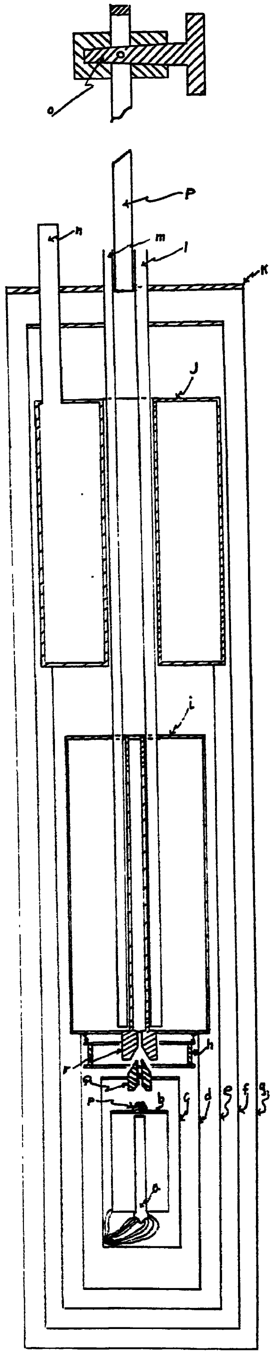

Fig. 7. Schematic cross section of assembled cryostat for low-temperature heat-capacity measurements.

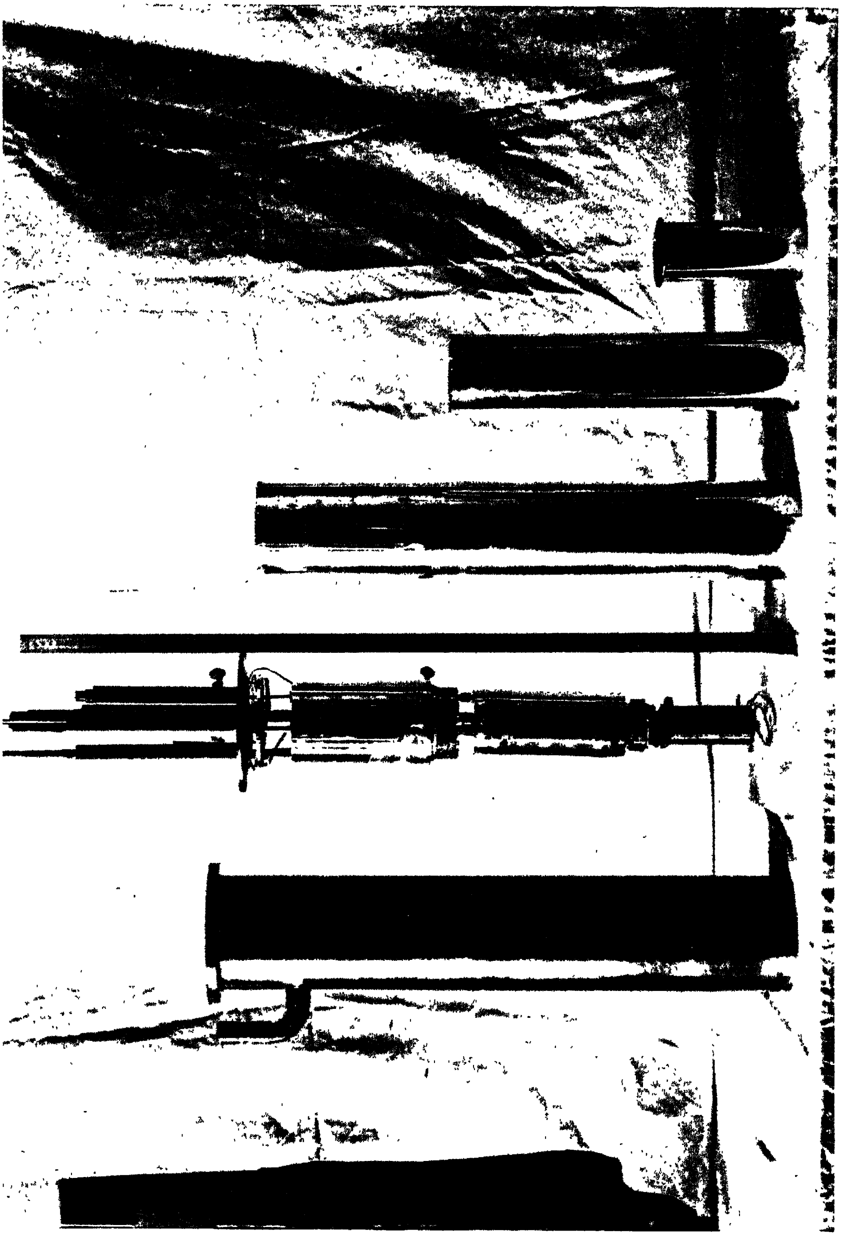

FIG. 8. Photograph showing parts of unassembled cryostat

constantan heater which covers nearly all of the external surface. The No. 32 enamel- and silk-covered copper lead wires are laid in a shallow helical groove, about 50 cm. in length, under the tube heater windings; good thermal contact and insulation for the heaters were assured by the application of electrician's enamel and baking. All of the leads are continuous to the outside and are thermally anchored successively to the floating ring (h), the hydrogen reservoir, and the liquid-air reservoir. The floating ring and the three parts of the adiabatic shield can be heated independently, and their relative temperatures may be observed by means of interconnected copper–constantan difference thermocouples in conjunction with a high-sensitivity Leeds & Northrup galvanometer. By this arrangement of heaters and thermocouples it is possible by suitable manual control to keep the ring, the shield, and the calorimeter at the same temperature, within close limits, and at any desired temperature above that of the hydrogen reservoir.[3]

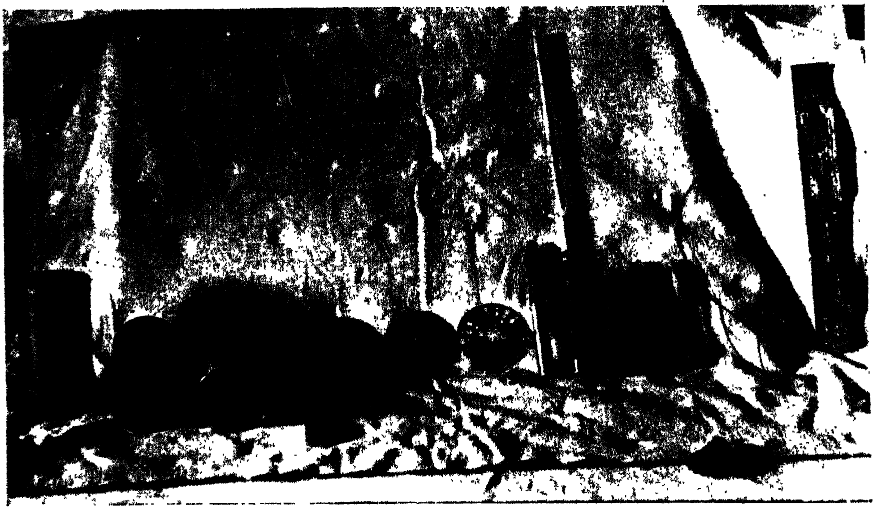

FIG. 9. Assembled calorimeter and parts shown at the left; assembled thermometer-heater and parts shown at the right.

The calorimeters are of copper and have the following approximate dimensions: length, 67 mm.; diameter, 35 mm.; internal volume, 55–60 ml.; mass, 45–50 g. Thin copper discs spaced along the reëntrant thermometer well and making contact with the wall aid in the rapid attainment of thermal equilibrium. A small screw and nut serve to secure one junction of the difference thermocouples to the bottom of the calorimeter. Liquids are admitted through a small tube in the top of the calorimeter, the tube being soldered shut after the remaining free space is filled with helium. The calorimeter parts and various stages of the assembly are shown in figure 9.

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[3] This method of maintaining adiabatic conditions in low-temperature calorimetry was first described by Southard and Brickwedde (J. Am. Chem. Soc. 55, 4378 (1933)).

THE METHOD OF HEAT TRANSFER

Heat transfer from the calorimeter and its radiation shield to the reservoirs of liquid air and liquid hydrogen is effected by a mechanical device or thermal switch, rather than by the more cumbersome method employing conduction through helium gas and the subsequent evacuation of the helium. In the present cryostat the calorimeter is attached to a stout cord (not shown) passing through a hole in the shield and terminating at a vacuum windlass (o), the windlass being on the outside of the brass can and hence always at room temperature. In operating the thermal switch the calorimeter and shield are drawn tightly into contact between accurately turned 60° conical fittings, shown at p, q, and r. When the temperature has dropped to that of the reservoir, the shield and calorimeter are lowered to their normal, isolated positions, a knot in the cord just above the shield serving to hold the calorimeter isolated in about the center of the adiabatic shield.

The cooling operation is normally carried out in two steps. In the first the calorimeter and shield are cooled to 51–52°K., both reservoirs in this operation containing liquid air under partial vacuum. In the second operation the liquid air in the lower reservoir is emptied and replaced by liquid hydrogen, which is subsequently frozen by pumping on it. The calorimeter and shield are again brought into contact with the hydrogen reservoir and held there until they have reached its temperature, this operation requiring 2–4 hr.

In general the cooling operations require a longer time than is the case when helium gas is used, but the advantages of the present method are obviously many. For example, a number of cooling and heating curves may be obtained in succession in the neighborhood of a transition region without having to break the vacuum in the apparatus.

TEMPERATURE SCALE AND RESISTANCE THERMOMETERS

The resistance thermometers

A set of five strain-free platinum resistance thermometers were constructed according to the design published by Myers (5). With one exception, each thermometer contained a coil of annealed platinum wire and a coil of No. 40 constantan wire mounted on a mica cross, the former serving for temperature measurements and the latter as a heater. The thermometer-heater assemblies were enclosed in 65 x 7.6 mm. platinum tubes, and the six platinum leads were brought out through glass heads sealed to the tubes. After repeated evacuation and annealing the thermometers were filled to $\frac{1}{2}$ atm. with purified helium and then sealed. The thermometer well in the bottom of the calorimeter is such that a thermometer fits snugly into it, thermal contact being made by the use of Apiezon grease M.

The thermometers were compared at frequent temperature intervals from 12–320°K. with one, H25, calibrated at the Bureau of Standards over the international temperature scale range, −183° to 444.6° C., and over the range 14–90°K. Carefully smoothed tables (4) of R/R_0 against temperature were constructed for both the standard thermometer and the working thermometer-heaters; as a result of the method used in calibrating the thermometers and in constructing the smoothed tables, the temperatures defined by all of the thermometers are internally self consistent over the entire range from 12–320°K. According to Brickwedde (3) the scale of the standard thermometer used by the Bureau may be expected to deviate as much as 0.02° from the thermodynamic scale in the range 14–90°K. In view of the accuracy with which R/R_0 may be measured, it is believed that both H25 and the working thermometer-heaters define the thermodynamic scale within 0.02 to 0.03° in the range 14–90°K. Above 90°K. the deviations may be larger than this, since it is known that the international scale does not coincide with the thermodynamic scale within the accuracy of measurements with resistance or gas thermometer. It will be evident from this discussion that the temperature scale consists of two parts: namely, the thermodynamic scale up to 90°K. and the international scale above 90°K.

Resistance measurements were made by comparing the drop in potential across the thermometer with that across a 100-ohm or 25-ohm Leeds & Northrup certified standard resistor in series with the thermometer. Below 36°K. a current of 4 ma. and the 25-ohm resistor were used; above 36°K. the current was 1 ma. and the 100-ohm standard was employed. A decade box in series with the thermometer permits adjustment of the current so that it is always the same no matter what the changes in the thermometer resistance or in the source of the current. By this arrangement the potentiometer reading with 1-ma. current gives a very close approximation of the resistance of the thermometer. The true values of the standards when used in resistance thermometry are not important, but a comparison with thermometer H25 at the ice point showed the standards to be within 0.01 per cent of the values certified by the manufacturer.

All potential measurements were made with a White autocalibrated double potentiometer and a sensitive galvanometer. A bank of six saturated cadmium cells certified by the Bureau of Standards serves as the laboratory standard of potential.

The first low-temperature calorimetric measurements were made at the Bureau in January 1945. Up to the time of this report (April 1946) measurements have been completed on sixteen compounds. One phase of this work is the study of complete series of isomers in order to determine the effect of branching on the entropy. With this point in view, studies have been made of the five hexanes and the six pentenes, and a study of the C_8H_{12} alkylcyclohexanes is in progress.

In addition to the above complete series, studies have been made on several other cycloparaffins.

Because of the importance of petroleum as a source material for many chemical compounds, the laboratory has started and will expand the work on sulfur and halogenated compounds.

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(1) Blanchard, E. R., and Bittner, H. W.: Rev. Sci. Instruments **13**, 394 (1942).
(2) Blue, R. W., and Hicks, J. F. G., Jr.: Rev. Sci. Instruments **59**, 1962 (1937).
(3) Brickwedde, F. G.: Private communication.
(4) Hoge, H. J., and Brickwedde, F. G.: Bur. Standards J. Research **22**, 351 (1939).
(5) Myers, C. H.: Bur. Standards J. Research **2**, 807 (1932).

THE USE OF ORGANOCADMIUM REAGENTS FOR THE PREPARATION OF KETONES

JAMES CASON

Department of Chemistry, University of California, Berkeley, California

Received August 26, 1946

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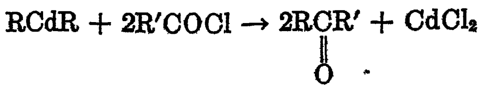
I. INTRODUCTION

The preparation of ketones by the reaction between an organocadmium compound and an acid chloride was first recommended by Gilman and Nelson (19) in 1936. The following reactions are involved in the preparation:

$$RBr + Mg \rightarrow RMgBr$$

$$2RMgBr + CdCl_2 \rightarrow RCdR + MgBr_2 + MgCl_2$$



$$RCdR + 2R'COCl \rightarrow 2R\underset{\underset{O}{\|}}{C}R' + CdCl_2$$

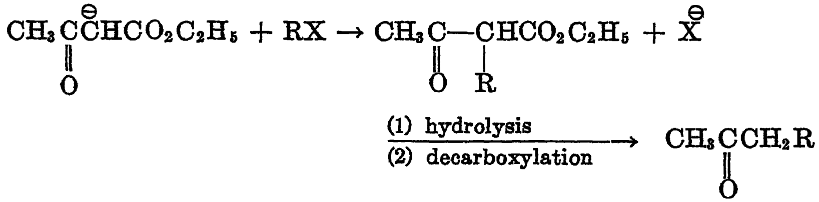
The cadmium reagent may be prepared from a lithium derivative instead of a Grignard reagent, if desired (19). This sequence may be carried through in a few hours without isolation of intermediates; however, the general utility of the method was not recognized for several years, possibly on account of experimental difficulties encountered in carrying out these reactions (*cf.* Section IV). At least one report (38) of a low yield in such a preparation has appeared. Since 1941, however, considerable information concerning this reaction has been gained, and the method may now be regarded as one of the best and most widely applicable procedures for laboratory-scale preparation of either simple ketones or various polyfunctional compounds containing a keto group. The required starting materials are relatively easily available, the experimental procedure is equivalent to a one-step process, and the yields are usually rather good.

II. Comparison with Other Methods

Although simple symmetrical ketones may be conveniently prepared from the appropriate acid or its salt, the most useful methods for preparing unsymmetrical ketones involve a Friedel–Crafts reaction, an enolate condensation, or some type of Grignard reaction. It is of interest, perhaps, to consider the relative merits of these methods as compared with the use of organocadmium reagents.

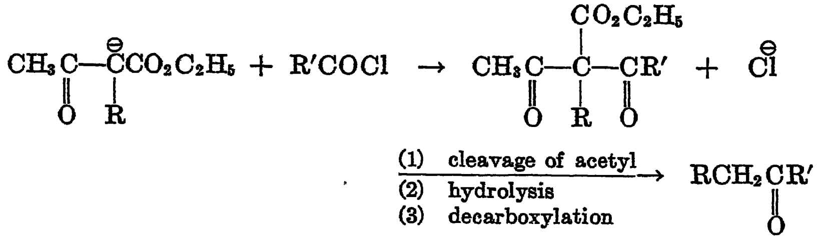
The Friedel–Crafts reaction is quite useful in cases where it is applicable. However, it is useful only in the aromatic series and with compounds containing no meta-directing groups; also, only certain orientations may be obtained. In the latter two categories mentioned above, the following general reactions may be considered:

(1)



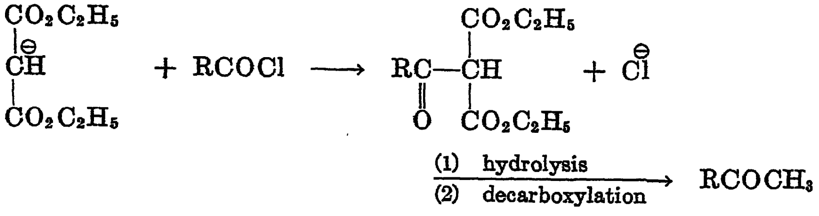
$$CH_3\underset{\substack{\|\\ O}}{C}\overset{\ominus}{C}HCO_2C_2H_5 + RX \rightarrow CH_3\underset{\substack{\|\\ O}}{C}-\underset{\substack{|\\ R}}{C}HCO_2C_2H_5 + \overset{\ominus}{X} \xrightarrow[\text{(2) decarboxylation}]{\text{(1) hydrolysis}} CH_3\underset{\substack{\|\\ O}}{C}CH_2R$$

(2)



$$CH_3\underset{\substack{\|\\ O}}{C}-\underset{\substack{|\\ R}}{\overset{\ominus}{C}}CO_2C_2H_5 + R'COCl \rightarrow CH_3\underset{\substack{\|\\ O}}{C}-\underset{\substack{|\\ R}}{\overset{\substack{CO_2C_2H_5\\ |}}{C}}-\underset{\substack{\|\\ O}}{C}R' + \overset{\ominus}{Cl} \xrightarrow[\substack{\text{(2) hydrolysis}\\ \text{(3) decarboxylation}}]{\text{(1) cleavage of acetyl}} RCH_2\underset{\substack{\|\\ O}}{C}R'$$

(3)



$$\underset{\substack{|\\ CO_2C_2H_5}}{\overset{\substack{CO_2C_2H_5\\ |\ominus}}{CH}} + RCOCl \longrightarrow R\underset{\substack{\|\\ O}}{C}-\underset{\substack{|\\ CO_2C_2H_5}}{\overset{\substack{CO_2C_2H_5\\ |}}{CH}} + \overset{\ominus}{Cl} \xrightarrow[\text{(2) decarboxylation}]{\text{(1) hydrolysis}} RCOCH_3$$

(4) $RMgX + R'C{\equiv}N \rightarrow \text{imine} \xrightarrow{\text{hydrolysis}} RCOR'$

(5) $RMgX + R'CONH_2 \rightarrow \text{imine} \xrightarrow{\text{hydrolysis}} RCOR'$

(6) $RMgX + (CH_3CO)_2O \xrightarrow{-70°\text{C.}} RCOCH_3 + CH_3CO_2MgX$

(7) $\underset{\substack{\text{or}\\ R_2Zn}}{RZnCl} + R'COCl \rightarrow RCOR' + ZnCl_2$

Method 1 has been widely used for the preparation of ketones; however, it involves two separate operations and the over-all yield of ketone rarely equals the yield obtained by use of the organocadmium reagent. This method may be extended to the preparation of keto acids (7, 32, 36) if a bromo ester is used instead of an alkyl halide; however, the difficult availability of most types of bromo esters is often a handicap. On the other hand, keto esters may be prepared in one step by the reaction of an organocadmium reagent with the ester acid chloride of a dibasic acid, $C_2H_5O_2C(CH_2)_nCOCl$, a type of starting material which is relatively easily obtained (11, 39). Method 1 has the advantage, however, that it may be used for the preparation of keto acids having a branching group between the keto and carboxyl groups.

The chief usefulness of method 1 is in the preparation of methyl ketones, for higher β-keto esters are obtainable (5, 30, 36) only by fairly laborious procedures. For this reason, method 2 has found considerable application in the preparation of other types of ketones, especially keto acids (32, 36). The chief difficulties with the method are that the over-all process is laborious and the yields are not especially good. Side reactions are oxygen-acylation and cleavage of the higher acyl group.

Method 3 is limited to the preparation of methyl ketones, for acylation of alkylmalonic esters (23, 40) gives a very poor yield, presumably owing to oxygen-acylation. For methyl ketones this method gives good yields, and it is about as useful as the cadmium method. For such ketones, there seems little basis for choice between the two methods.

Methods 4 (2, 17, 18) and 5 (17, 22, 38) give good yields of ketones, provided there is present in the molecule no other functional group which will react with the Grignard reagent. The nitrile and amide groups are among the least reactive toward this reagent. If this condition is met these methods give yields as good as or better than the cadmium method, and the experimental manipulations are similar. The scope of the cadmium method is much greater however, for the organocadmium reagents fail to react with most functional groups (*cf.* Section III,D).

Method 6 has been introduced only recently (27, 28) but seems quite promising as a general preparative method for methyl ketones. The yields are in the same range as those obtained from cadmium reagents. It is also reported that this method is satisfactory with secondary and tertiary alkyl halides, types which are not suitable for conversion to dialkylcadmium reagents.

Method 7 is identical with the preparation involving cadmium, except that zinc is used. Although Blaise (3) originally claimed excellent yields in this type of preparation, a succession of investigators (25, 26, 33, 34) has failed to obtain better than moderate yields. The superiority of the cadmium reagent arises from its much easier preparation and its lower reactivity toward the carbonyl group. Diethylzinc has been shown (19) to react with the carbonyl group about 3.5 times as fast as does diethylcadmium. A few ketones have been made from both cadmium and zinc reagents (15, 19), and in each case the yield was significantly larger when the cadmium reagent was used.

Thus, it seems safe to say that no other method of ketone preparation has the scope of the method utilizing the organocadmium reagents, and in relatively few specific cases are other methods superior. The actual scope and limitations of this method will be considered next.

III. Scope and Limitations of the Method

A. THE ALKYL OR ARYL HALIDE

1. The halogen

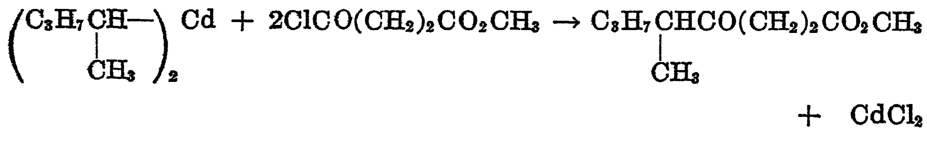
Several investigators (12, 16, 19) have reported that if the Grignard reagent from which the cadmium derivative is prepared is made from an iodide, the yield of ketone is much poorer than if the bromide is used. It has also been reported (12) that in the case of *n*-butyl halides the chloride is somewhat inferior to the bromide, although much better than the iodide. In the preparation of methyl 4-ketoöctoate from di-*n*-butylcadmium and β-carbomethoxypropionyl chloride, the yields were respectively 79.5, 63, and 45 per cent when the halides

$$(C_4H_9)_2Cd + 2ClCO(CH_2)_2CO_2CH_3 \rightarrow C_4H_9CO(CH_2)_2CO_2CH_3 + CdCl_2$$

were butyl bromide, chloride, and iodide. Cole and Julian (15) reported the preparation of a methyl ketone using dimethylcadmium, and obtained the same yield when either methyl iodide or methyl bromide was used; however, a fivefold excess of cadmium reagent was used in this preparation, so the relative value of the halides may have been obscured. This seems probable, since De Benneville (16) reported methyl iodide inferior to the bromide.

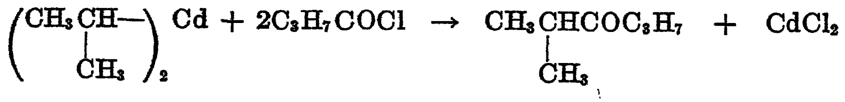
2. The organic radical

Aromatic Grignard reagents form cadmium derivatives readily (12, 15, 16, 19) and give good yields of aromatic ketones. If the organic radical is alkyl, then it must be primary in order for the preparation to be useful. This seems to be the chief limitation of the organocadmium reagents. Gilman and Nelson (19) first reported that secondary and tertiary cadmium derivatives are stable only at low temperature. Other investigators (12, 14) have found that in the preparation of methyl 4-keto-5-methyloctoate from di-2-amylcadmium and β-carbomethoxypropionyl chloride, the yield of keto ester is nearly zero if the reaction



$$\left(\underset{\displaystyle CH_3}{C_3H_7\overset{}{C}H}-\right)_2 Cd + 2ClCO(CH_2)_2CO_2CH_3 \rightarrow C_3H_7\underset{\displaystyle CH_3}{CH}CO(CH_2)_2CO_2CH_3 + CdCl_2$$

temperature is allowed to rise somewhat above 0°C. (12a), and the yield is only 21.5 per cent when the reaction is conducted at $-5°$ to $-10°C$. Diisopropylcadmium seems to be somewhat more stable, for Gilman and Nelson, operating at 0°C., obtained *n*-propyl isopropyl ketone in 60 per cent yield:



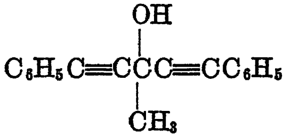
$$\left(\underset{\displaystyle CH_3}{CH_3CH}-\right)_2 Cd + 2C_3H_7COCl \rightarrow \underset{\displaystyle CH_3}{CH_3CH}COC_3H_7 + CdCl_2$$

Di-*tert*-butylcadmium has been used in the reaction with acetyl chloride (19) at −70°C. to give a 17 per cent yield of ketone. De Benneville (16) reported a 40 per cent yield (based on anhydride) in the reaction, at −70°C., between di-*tert*-butylcadmium and benzoic anhydride; however, in view of the low reactivity of cadmium reagents toward anhydrides (*cf.* Section III,C) and the preparation of ketones from Grignard reagents and anhydrides at −70°C. (27), it seems quite possible that the ketone obtained in this case resulted from Grignard reagent rather than cadmium reagent.

It may be mentioned at this point that Nightingale and Wadsworth (29) attempted the conversion of phenylethynylmagnesium bromide (I) to the corresponding cadmium derivative. After stirring for 1 hr. in ether the test for Grignard reagent was still positive, and when the resultant mixture was treated with acetic anhydride or acetyl chloride the only product isolable (in 14 per cent yield) was that expected from the Grignard reagent, bis(phenylethynyl)methylcarbinol (II).

$$C_6H_5C{\equiv}CMgBr$$

I



$$C_6H_5C{\equiv}C\overset{\displaystyle OH}{\underset{\displaystyle CH_3}{C}}C{\equiv}CC_6H_5$$

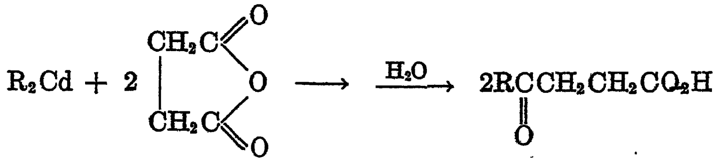
II

B. THE CADMIUM HALIDE

It has been shown (19) that cadmium chloride is at least as effective as is cadmium bromide for preparing the cadmium reagents. Since the bromide is more expensive and much more hygroscopic, the chloride is always used. It has also been reported (19) that the yields are approximately the same from a dialkylcadmium compound and an alkylcadmium halide; so the former is customarily used, since it requires only half as much cadmium chloride.

C. THE ACID CHLORIDE OR ANHYDRIDE

Organocadmium reagents will react with acid anhydrides (10, 12, 14, 16); however, the reaction seems to be uniformly inferior to that obtained with acid chlorides. The most interesting reaction with anhydrides is that with cyclic anhydrides. Unfortunately, the yields of keto acids (on the basis of alkyl or aryl halide) in this reaction have been only 25–30 per cent; so unless the halide is quite cheap, it seems profitable to convert the anhydride to the half-ester and this to the ester acid chloride (11, 39). Reaction of the latter compound with the cadmium reagent gives the keto ester corresponding to the keto acid obtained



from the anhydride (*cf.* equation on page 18). An exception is the phthalic anhydride type, for the ester acid chlorides of such dibasic acids are unstable (45).

Both aliphatic and aromatic acid chlorides react smoothly, the aliphatic ones being more reactive. This preparation of ketones has been very useful in the sterol series (15, 21, 21a, 29a), and a good yield has been obtained from a substituted acid chloride in the cyclohexane series (23). An aliphatic acid chloride of as high molecular weight as stearyl chloride has been found to react satisfactorily (19, 35). A considerable number (*cf.* table 2) of keto esters have been obtained in good yield from ester acid chlorides of dibasic acids. There has been one report (29) of the use of α,β-unsaturated acid chlorides in the reaction with an organocadmium reagent. The reaction between cinnamoyl chloride and diphenylcadmium gave about as good a yield as normal for the conditions used; however, several dialkylcadmium reagents gave only viscous oils in the reaction with cinnamoyl chloride, crotonyl chloride, or phenylpropiolyl chloride.

Bunnett and Tarbell (8) recently reported the interesting synthesis of α-chloro ketones by the reaction of organocadmium compounds with α-chloro acid chlorides, and more recently (12) the yields in such preparations have been improved considerably. The process is illustrated by the synthesis of 1-chloro-2-hexanone (III):

$$(C_4H_9)_2Cd + 2ClCH_2COCl \rightarrow \underset{\substack{\text{III}\\ \text{1-Chloro-2-hexanone}}}{2C_4H_9COCH_2Cl} + CdCl_2$$

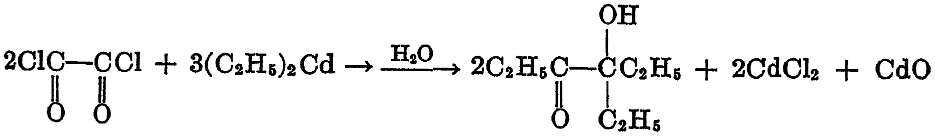
This preparation of α-chloro ketones is much simpler and safer than that involving an intermediate diazoketone (1, 4), especially if considerable quantities of such ketones are desired.

D. REACTION OF CADMIUM REAGENTS WITH OTHER GROUPS

1. *Carbonyl groups*

An isolated carbonyl group reacts with the cadmium reagent, but very slowly indeed. After 0.03 mole of benzaldehyde had been allowed to stand in a sealed tube for five months with excess diethylcadmium, there was obtained only 1.3 g. of ethylphenylcarbinol (19). Thus, the preparation of ketones by adding the acid chloride to the cadmium reagent results in very little addition of the cadmium reagent to the carbonyl group. It is true, however, that in some cases (12) the yield is improved somewhat by making the above addition at a low temperature and warming the mixture after all the acid chloride has been added.

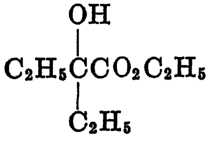
If the carbonyl group is activated by an adjacent group such as an ester group, the cadmium reagent may add to the activated carbonyl. For example (19), the reaction between oxalyl chloride and diethylcadmium gives a 27 per cent yield of diethylpropionylcarbinol:



$$2Cl\underset{\underset{O}{\|}}{C}-\underset{\underset{O}{\|}}{C}Cl + 3(C_2H_5)_2Cd \xrightarrow{H_2O} 2C_2H_5\underset{\underset{O}{\|}}{C}-\overset{\overset{OH}{|}}{\underset{\underset{C_2H_5}{|}}{C}}C_2H_5 + 2CdCl_2 + CdO$$

2. Ester groups

Diethylcadmium reacts with diethyl oxalate to give ethyl α-ethyl-α-hydroxybutyrate (IV) (19); however, cadmium reagents appear not to react at an appreciable rate with isolated ester groups, even formates (21). In the preparation of ethyl 10-ketohendecanoate from dimethylcadmium and ω-carbethoxynonoyl chloride (14) there was isolated a small amount of 2,11-dodecanedione (V), and it was suggested that this may have arisen from some reaction of dimethylcadmium with the ester grouping in ethyl 10-ketohendecanoate. This now seems highly improbable, and the trace of diketone very likely resulted from reaction of the dimethylcadmium with a small amount of sebacyl dichloride (VI) present in the ester acid chloride. The ester acid chloride was made from ethyl hydrogen sebacate, and half-esters disproportionate to some extent, on long standing, to the diacid and diester (12).



$$C_2H_5\overset{\displaystyle OH}{\underset{\displaystyle C_2H_5}{C}}CO_2C_2H_5$$

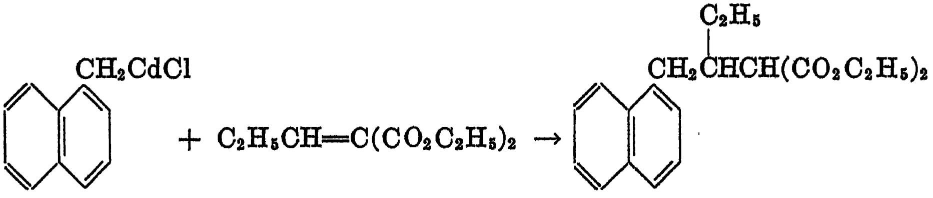
IV

$$(CH_3)_2Cd + 2ClCO(CH_2)_8CO_2C_2H_5 \rightarrow 2CH_3CO(CH_2)_8CO_2C_2H_5 + CdCl_2$$

Ethyl 10-ketohendecanoate

$CH_3CO(CH_2)_8COCH_3$	$ClCO(CH_2)_8COCl$
V	VI
2,11-Dodecanedione	Sebacyl dichloride

It has recently been reported (31) that cadmium reagents add in a 1,4-manner to alkylidenemalonic esters. In fact, the yields obtained in several such reactions were better than those obtained when the corresponding Grignard reagent was used. A representative reaction is the following, which gave a 65 per cent yield:



3. Miscellaneous groups

Only limited data are available concerning the reaction of cadmium reagents with other groups than those mentioned; however, it may be safely assumed that this reagent will not add to an amide or nitrile grouping. Gilman and Nelson (19) investigated several groupings and found that (*a*) coupling with *tert*-butyl chloride did not occur; (*b*) reaction with phenyl isocyanate did not occur (the isocyanate polymerized); (*c*) reaction with benzenesulfonyl chloride was according to the following equation:

Coupling does occur readily (37) with the α-chloro ether, 2,3-dichloro-1,4-dioxane, and the yield of 2,3-dialkyl-1,4-dioxane is much better than when the Grignard reagent is used. There seems to be no report concerning the reactivity of cadmium reagents toward the aromatic nitro grouping.

From the data in this section it is seen that, with the exception of groups containing active hydrogen, very few groups interfere with the preparation of ketones from cadmium reagents.

E. SIDE REACTIONS

1. *Ester formation*

It is known that acid chlorides will react with diethyl ether in the presence of anhydrous magnesium halide (10, 41, 42):

$$RCOCl + (C_2H_5)_2O \xrightarrow{Mg^{++}} RCO_2C_2H_5 + C_2H_5Cl$$

Since there is present in the mixture during a cadmium reaction a full molar equivalent of magnesium halide, the conditions for reaction of the acid chloride with ether are good. Although Gilman and Nelson (19) indicated that this reaction occurs to a very small extent, later workers (10, 12, 14, 29a) obtained appreciable quantities of ester when the reaction was carried out in ether. If the ether solution of the cadmium reagent is distilled to dryness and the reaction with the acid chloride is carried out in benzene solution, ester formation is greatly reduced, but there is still obtained 0.5–2.5 per cent of ester. This is true even if, after distillation of ether to dryness, benzene is added and distilled.

This small amount of ester is ordinarily of no great significance. It may be separated, if desired, by fractional distillation unless the cadmium reagent contains methyl or ethyl radicals. Methyl and ethyl ketones boil very near the corresponding ethyl esters. In such cases, a pure sample of ketone may be obtained by saponification and removal of the resultant acid.

If butyl ether is used as solvent (12) for preparation of the cadmium reagent, ester formation is reduced by the lower reactivity of butyl ether; also, the resulting butyl ester may be separated by fractional distillation from the corresponding methyl or ethyl ketones. Unfortunately, these advantages are nullified by the fact that the yields obtained in butyl ether are low (in the case reported), and Grignard reagents react very sluggishly with cadmium chloride in butyl ether. In the preparation of ethyl 10-ketohendecanoate (*cf.* page 21) in benzene as solvent the yield was 84 per cent, whereas in butyl ether it was only 57 per cent.

2. *Addition to carbonyl*

As stated in Section III,D,1, this reaction occurs, but to only a small extent unless the carbonyl is activated.

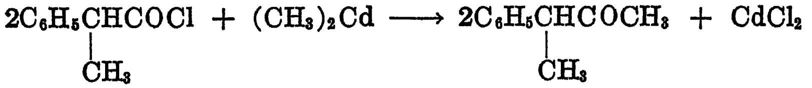
3. *Enolization*

The reaction of the cadmium reagent with the enol form of the ketone being prepared is always a potential side reaction. Gilman and Nelson (19) showed that benzaldehyde may be condensed with acetophenone in the presence of diethylcadmium:

$$C_6H_5CHO + CH_3COC_6H_5 \xrightarrow[\text{21 per cent yield}]{(C_2H_5)_2Cd} C_6H_5CH{=}CHCOC_6H_5$$

It was suggested that an —OCd type of compound might be involved in this reaction.

For most ketones it seems apparent that reaction of the enol form with cadmium reagent is a very minor side reaction. In fact, Campbell and Kenyon (9) carried out the following reaction, using *d*-methylphenylacetyl chloride:



$$2C_6H_5\underset{\substack{|\\ CH_3}}{C}HCOCl + (CH_3)_2Cd \longrightarrow 2C_6H_5\underset{\substack{|\\ CH_3}}{C}HCOCH_3 + CdCl_2$$

Since no significant racemization occurred during this reaction, it may be assumed that the concentration of the enol form of the ketone was very small. It should be pointed out that retention of asymmetry in the ketone may have resulted from the fact that nearly all the enol formed during the reaction was lost by condensation with another molecule of ketone. This is consistent with the fact that the good yields obtained in the preparation of simple ketones indicate very little loss of cadmium derivative by enolization.

In the case of α-chloro ketones, the loss by enolization becomes rather serious (12), so that the yields obtained in this type of preparation are lower than normal, 40–60 per cent. This, presumably, is due to the presence of the halogen atom in the α-position, promoting enolization. The occurrence of this reaction during the preparation of α-chloro ketones was established by isolation and characterization of the hydrocarbon formed by reaction of the cadmium derivative with the enol. When the preparation of 2-chloro-3-heptanone was carried out in ether,

$$(C_4H_9)_2Cd + 2CH_3CHClCOCl \longrightarrow \underset{\text{2-Chloro-3-heptanone}}{2C_4H_9COCHClCH_3} + CdCl_2$$

the yield was lower than when the preparation was carried out in benzene, and the yield of butane was greater in the case where the yield of ketone was lower. This indicates that enolization is more serious in ether as solvent than in benzene.

It was also found (12) that 1-chloro-2-hexanone (III) reacts rapidly and exothermically with di-*n*-butylcadmium, and there was evolved an amount of butane which, after purification, was equivalent to 70 per cent of the butyl bromide used for preparing the dibutylcadmium. When this reaction was worked up, very little of the 1-chloro-2-hexanone was recovered, most of it having been converted to higher-boiling materials. This could result from condensation of the cadmium enolate with another molecule of the chloro ketone.

IV. EXPERIMENTAL CONSIDERATIONS

Yields in the preparation of ketones from cadmium reagents can be greatly reduced by an incorrect experimental procedure, and it is probably in order to make some reference to the critical study of these factors which has been made (12). Certain general factors concerning experimental procedures were mentioned in Section III.

TABLE 1

Simple ketones and α-chloro ketones

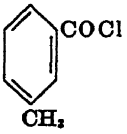
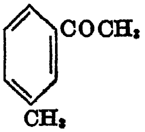
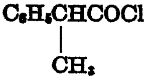
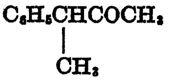
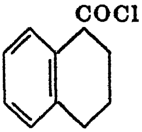
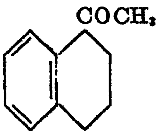
HALIDE	ACID CHLORIDE OR ACID ANHYDRIDE	KETONE	YIELD	REFERENCE
			per cent	
C_2H_5Br	CH_3COCl	$C_2H_5COCH_3$	50	(19)
n-C_4H_9Br	CH_3COCl	$C_4H_9COCH_3$	74	(19)
n-C_4H_9Br	$(CH_3CO)_2O$	$C_4H_9COCH_3$	56	(16)
$(CH_3)_3CBr$	CH_3COCl	$(CH_3)_3CCOCH_3$	17	(19)
$(CH_3)_2CHBr$	C_3H_7COCl	$(CH_3)_2CHCOC_3H_7$	60	(19)
C_6H_5Br	CH_3COCl	$C_6H_5COCH_3$	83	(19)
CH_3Br	C_6H_5COCl	$C_6H_5COCH_3$	85	(19)
C_6H_5Br	$(CH_3CO)_2O$	$C_6H_5COCH_3$	75	(16)
C_6H_5Br	C_2H_5COCl	$C_6H_5COC_2H_5$	81	(12, 19)
C_2H_5Br	C_6H_5COCl	$C_6H_5COC_2H_5$	84	(12, 19, 38)
C_6H_5Br	$(C_2H_5CO)_2O$	$C_6H_5COC_2H_5$	68	(16)
C_2H_5Br	$(C_6H_5CO)_2O$	$C_6H_5COC_2H_5$	53	(16)
$C_6H_5CH_2Cl$	CH_3COCl	$C_6H_5CH_2COCH_3$	18	(19)
CH_3Br			83	(19)
CH_3Br	 $C_6H_5\underset{\underset{CH_3}{\vert}}{C}HCOCl$	 $C_6H_5\underset{\underset{CH_3}{\vert}}{C}HCOCH_3$	78	(9)
C_6H_5Br	$((CH_3)_2CHCO)_2O$	$C_6H_5COCH(CH_3)_2$	72	(16)
$(CH_3)_2CHBr$	$(C_6H_5CO)_2O$	$C_6H_5COCH(CH_3)_2$	44	(16)
CH_3Br			65	(28)

TABLE 1—*Continued*

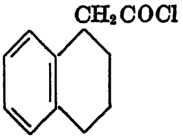
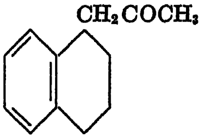
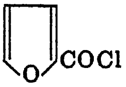
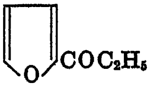
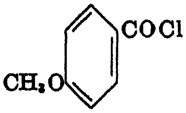
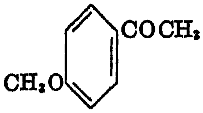
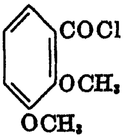
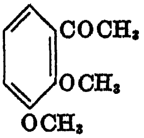
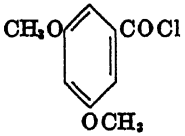
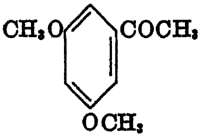
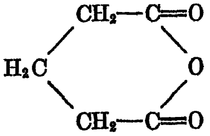
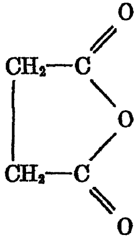
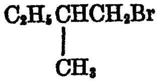
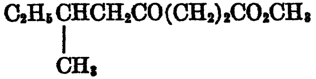
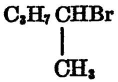
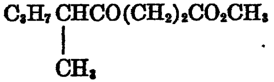
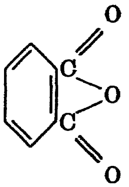
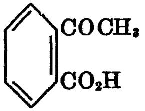
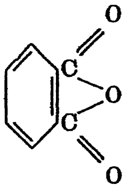
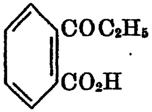
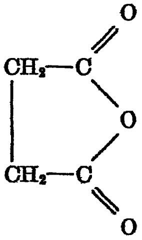
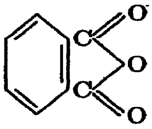
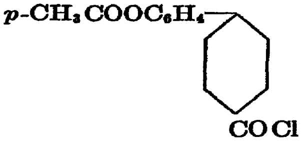
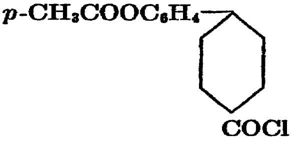
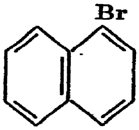
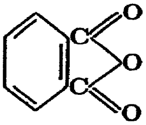
HALIDE	ACID CHLORIDE OR ACID ANHYDRIDE	KETONE	YIELD	REFERENCE
			per cent	
CH_3Br			69	(28)
C_6H_5Br	C_6H_5COCl	$C_6H_5COC_6H_5$	57	(19)
C_6H_5Br	$C_6H_5CH{=}CHCOCl$	$C_6H_5CH{=}CHCOC_6H_5$	44	(29)
C_2H_5Br	n-$C_{17}H_{35}COCl$	$C_{17}H_{35}COC_2H_5$	65	(19)
$C_6H_5(CH_2)_2Br$	n-$C_{17}H_{35}COCl$	$C_{17}H_{35}CO(CH_2)_2C_6H_5$	65.5	(35)
n-C_4H_9Br	$CH_2ClCOCl$	$C_4H_9COCH_2Cl$	63.5	(8, 12)
n-$C_5H_{11}Br$	$CH_2ClCOCl$	$C_5H_{11}COCH_2Cl$	24	(8)
n-C_4H_9Br	$CH_3CHClCOCl$	$C_4H_9COCHClCH_3$	53	(12)
n-$C_{12}H_{25}Br$	$CH_2ClCOCl$	$C_{12}H_{25}COCH_2Cl$	18	(8)
n-$C_{12}H_{25}Br$	$CH_3CHClCOCl$	$C_{12}H_{25}COCHClCH_3$	13	(8)
n-$C_{14}H_{29}Br$	$CH_2ClCOCl$	$C_{14}H_{29}COCH_2Cl$	41	(8, 12)
C_2H_5Br			61	(19)
CH_3Br			84	(19)
CH_3Br			71	(44)
CH_3Br			84	(44)
CH_3Br	$C_6H_5O(CH_2)_3COCl$	$C_6H_5O(CH_2)_3COCH_3$	78	(6)

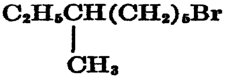
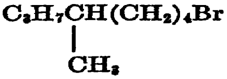
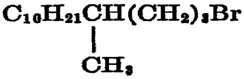
TABLE 2

Keto acids and keto esters

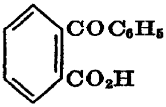
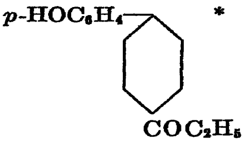
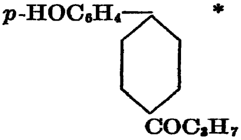
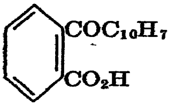
HALIDE	ACID CHLORIDE OR ACID ANHYDRIDE	KETONE	YIELD	REFERENCE
			per cent	
CH_3Br	$ClCO(CH_2)_4CO_2C_2H_5$	$CH_3CO(CH_2)_4CO_2C_2H_5$	76	(14, 24)
C_4H_9Br	$ClCO(CH_2)_2CO_2CH_3$	$C_4H_9CO(CH_2)_2CO_2CH_3$	80	(12)
n-C_4H_9Br		$C_4H_9CO(CH_2)_3CO_2H$	30.5	(12)
$(CH_3)_2CH(CH_2)_2Br$	$ClCO(CH_2)_2CO_2CH_3$	$(CH_3)_2CH(CH_2)_2CO(CH_2)_2CO_2CH_3$	79	(10, 12, 14)
$(CH_3)_2CH(CH_2)_2Br$		$(CH_3)_2CH(CH_2)_2CO(CH_2)_2CO_2H$	31	(14)
 $C_2H_5\underset{\overset{\vert}{CH_3}}{CH}CH_2Br$	$ClCO(CH_2)_2CO_2CH_3$	 $C_2H_5\underset{\overset{\vert}{CH_3}}{CH}CH_2CO(CH_2)_2CO_2CH_3$	60	(14)
 $C_3H_7\underset{\overset{\vert}{CH_3}}{CH}Br$	$ClCO(CH_2)_2CO_2CH_3$	 $C_3H_7\underset{\overset{\vert}{CH_3}}{CH}CO(CH_2)_2CO_2CH_3$	21.5	(12, 14)

CH_3Br			62	(16)
C_2H_5Br			67	(16)
C_6H_5Br		$C_6H_5CO(CH_2)_2CO_2H$	30	(16)
CH_3Br	$ClCO(CH_2)_7CO_2C_2H_5$	$CH_3CO(CH_2)_7CO_2C_2H_5$	64	(24)
CH_3Br	$ClCO(CH_2)_8CO_2C_2H_5$	$CH_3CO(CH_2)_8CO_2C_2H_5$	84	(12, 14)
$(CH_3)_2CH(CH_2)_2Br$	$ClCO(CH_2)_4CO_2CH_3$	$(CH_3)_2CH(CH_2)_2CO(CH_2)_4CO_2CH_3$	84	(43)

HALIDE	ACID CHLORIDE OR ACID ANHYDRIDE
C_6H_5Br	
C_2H_5Br	
n-C_3H_7Br	
$(CH_3)_2CH(CH_2)_2Br$	$ClCO(CH_2)_8CO_2C_2H_5$
	
$(CH_3)_2CH(CH_2)_6Br$	$ClCO(CH_2)_8CO_2C_2H_5$

$C_2H_5\underset{\underset{\displaystyle CH_3}{\vert}}{CH}(CH_2)_5Br$	$ClCO(CH_2)_8CO_2C_2H_5$
$C_3H_7\underset{\underset{\displaystyle CH_3}{\vert}}{CH}(CH_2)_4Br$	$ClCO(CH_2)_8CO_2C_2H_5$
$C_{10}H_{21}\underset{\underset{\displaystyle CH_3}{\vert}}{CH}(CH_2)_3Br$	$ClCO(CH_2)_8CO_2C_2H_5$
C_6H_5Br	

2 — *Continued*

KETONE	YIELD *per cent*	REFERENCE
	64	(16)
	63	(23)
	75	(23)
$(CH_3)_2CH(CH_2)_2CO(CH_2)_8CO_2C_2H_5$	85	(14)
	57	(16)
$(CH_3)_2CH(CH_2)_6CO(CH_2)_8CO_2C_2H_5$	46	(10)

* * *

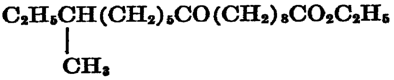
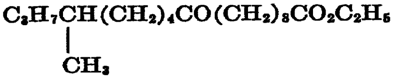
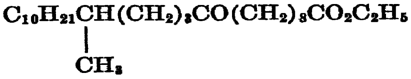
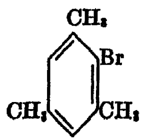
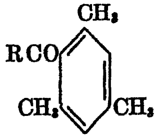
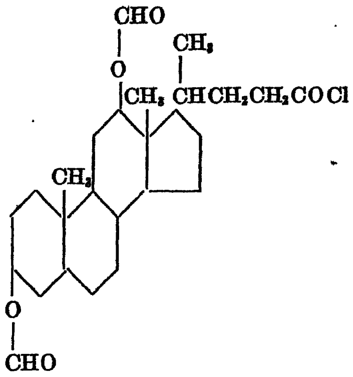
$C_2H_5\underset{CH_3}{CH}(CH_2)_5CO(CH_2)_8CO_2C_2H_5$	76.5	(14)
$C_3H_7\underset{CH_3}{CH}(CH_2)_4CO(CH_2)_8CO_2C_2H_5$	65	(13)
$C_{10}H_{21}\underset{CH_3}{CH}(CH_2)_3CO(CH_2)_8CO_2C_2H_5$	78	(13)
$RCOC_6H_5$ (R = the sterol nucleus indicated)	98	(15)

TABLE 2—*Continued*

HALIDE	ACID CHLORIDE OR ACID ANHYDRIDE	KETONE	YIELD	REFERENCE
			per cent	
	RCOCl		46	(15)
$(CH_3)_2CHBr$	RCH_2CH_2COCl	$RCH_2CH_2COCH(CH_3)_2$ (R = sterol nucleus indicated above)		(29a)
C_6H_5Br		$R'COC_6H_5$† (R′ = sterol nucleus indicated at left, except that the formyl esters at the 3- and 12-positions were hydrolyzed to give hydroxyl groups)	81	(21)

---

* The keto ester obtained as the initial reaction product was hydrolyzed and this product was isolated.

† By the same process, phenyl ketones of seven other steroid derivatives were prepared by these authors. This preparation, only, was described "as an example", and yields were not reported for the other preparations. Jacobsen (21a) has also reported a similar preparation.

After the Grignard reagent has been prepared in ether it is allowed to react with solid cadmium chloride in this solvent, most conveniently at the boiling temperature unless a thermally unstable cadmium derivative is being prepared. The length of time required for this reaction varies considerably, even for the same preparation, probably owing to variation in particle size of the cadmium chloride. Thus, completion of this reaction should always be checked by testing for Grignard reagent (12, 19).

After formation of the cadmium derivative, it is desirable to remove ether by distillation and carry out the reaction with the acid chloride in benzene. This nearly eliminates ester formation (*cf.* Section III,E,1), reduces enolization (*cf.* Section III,E,3), and often has the additional advantage of preventing the reaction complex from agglutinating and stopping the stirrer, in cases where this happens in ether. Also, the higher reaction temperature makes possible the completion of the reaction in 10–60 min., whereas several hours are sometimes required in boiling ether.

The laborious procedures (6, 8, 37) for drying cadmium chloride that have been described are quite unnecessary. Drying to constant weight in an oven at 110°C. gives entirely reliable anhydrous cadmium chloride. It may be handled in air without absorbing significant amounts of moisture; however, it should be stored in a desiccator, for it slowly absorbs moisture from the air.

The organometallic compounds and metallic halides formed in this reaction always precipitate and form heavy sludges. Failure to continue adequate stirring to the end of the reaction lowers the yield significantly; so an efficient stirrer, such as the Hershberg type (20), and a powerful stirring motor should be used.

Since there is some loss in preparing the cadmium derivative from the alkyl or aryl halide *via* the Grignard reagent, there is no advantage in using an amount of acid chloride equivalent to the bromide. It has been found (12) that the maximum useful ratio of acid chloride to bromide is about 0.8. Use of a larger ratio of acid chloride does not increase the yield based on bromide. If methyl or ethyl bromide is used, some loss of dialkylcadmium occurs when the ether is distilled from the mixture, so it is advisable to use only a 0.5 molar equivalent of acid chloride if complete consumption of the acid chloride is to be assured. In cases where a very expensive acid chloride is used, a large excess of cadmium reagent is recommended for a maximum yield of ketone based on the acid chloride. In this manner, yields up to 98 per cent have been obtained (15).

V. KETONES PREPARED BY THE CADMIUM REACTION

An effort has been made to enter in tables 1 and 2 all the ketones prepared by the reaction of an organocadmium reagent with an acid chloride or acid anhydride; however, it is probable that isolated examples reported as part of a synthesis have been overlooked. Since the factors contributing to a maximum yield in this type of preparation have been investigated only recently, it is certain that many of the yields included in the tables lack much of being the maximum obtainable. The yields in the tables are based on acid chloride or acid anhydride.

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THE CHEMICAL COMPOSITION OF BARKS[1]

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Bark analyses and components are tabulated and discussed. The standard wood analytical procedures for lignin and cellulose when applied to barks are often inadequate. The lignin fraction from barks usually contains materials low in methoxyl and, in some instances, carboxyl groups. There is evidence that the carbohydrate fraction in certain barks is chemically bound to the lignin fraction by bonds absent in wood. Barks contain hemicelluloses and cellulose; the sugars associated with the hemicelluloses are glucose, galactose, mannose, arabinose, and xylose, whereas the cellulose may have a large percentage of the alkali-resistant alpha cellulose. Chemically, little is known about the cork in barks. Numerous compounds have been isolated from bark extractives. These include representatives of the alkaloids, sterols and various alcohols, esters, glucosides, hydrocarbons, mucilages, proteins, saponins, sugars, tannins, and many types of acids and pigments. Ash, nitrogen, and the B.T.U. contents of barks are higher than for wood.

The literature on barks has been reviewed to the end of 1945.

The trunk of a mature tree is composed of a core of wood covered by a layer of bark on the outside. Inasmuch as the bark lies outside of the living cambium, it is pushed outward by the development of new tissue underneath. This increase in circumference to which the bark must accommodate itself often exceeds its ability to respond, with the result that the older tissues are broken and killed by exposure, and especially by the stoppage of food and water supplies. Thus two bark layers may be recognized: (*a*) the outer, dark-colored, dead portion that varies greatly in thickness, and (*b*) the inner, thin, light-colored, living portion. A roughening of the bark is held to indicate the inception of cork formation. Bark tissues may contain several types of cells, often designated as sieve, longitudinal parenchyma, ray parenchyma, stone, cork, fiber, etc., which may contribute to their different properties.

On the whole the available literature on barks gives too little information concerning their chemical composition, and much less is known of the properties of the actual components than one would normally expect. Millions of tons of bark from the lumber and pulp industries are wasted or destroyed annually. Wilson (80) has estimated that 209,000,000 cu. ft. of bark were developed in the manufacture of about 10.6 billion board feet of lumber at primary sawmilling operations in 1944 in the states of Oregon and Washington alone. About 12.5–14.7 per cent of the sawlogs in these states is bark. On the other hand, the chemical utilization of bark from cut timber is a field of research which has hardly been touched. Newer methods of bark removal from logs with the hydraulic barker give cleaner bark and will effect economies. It would appear that the availability of such large amounts of a natural raw material would stimulate a

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great deal of research activity in the hope that such work will give rise to fuller utilization. In many instances the information needed for evaluation of barks having possible technological interest could be furnished by appropriate chemical analyses if these were available. In this connection the investigations of Zellner (81, 82, 83) and more recently of Sharkov (60–68) are particularly significant.

BARK ANALYSES

Our knowledge of the composition of barks is due preëminently to the application of analytical methods used for wood. That these procedures have not been entirely satisfactory when applied to barks is evident from the absence of cellulose and lignin determinations in table 1. No lignin determinations are reported for most barks, and what would ordinarily be reported as cellulose in wood analyses is designated as fiber because of questionable delignification. Table 2 lists extraneous organic compounds that were found to occur in the individual barks.

In the light of the information contained in tables 1 and 2, the principal components of bark with which the chemist is concerned are:

1. Lignin—the material insoluble in concentrated mineral acids
2. Cork—cutose, suberin, and suberic acid (1,6-hexanedicarboxylic acid)
3. Carbohydrates—holocellulose, the total carbohydrate fraction
   A. Cellulose
   B. Hemicelluloses—arabans, xylans, mannans, glucosans, galactans, and uronic acid substances
4. Extraneous materials
   A. Volatile acids and oils
   B. Non-volatile fatty oils (fats and fatty acids), higher alcohols, resins, and hydrocarbons
   C. Coloring matters
   D. Tannins and the related water-insoluble phlobaphenes
   E. Polysaccharides, glucosides, pectins, and sugars
   F. Organic nitrogen compounds
   G. Mineral matters
   H. Other organic components—saponins, mannitol, dulcitol, etc.

It can be readily understood that this classification indicates only types of the principal components present in barks and that in all cases, when not dealing with crystalline compounds, these components are likely to be of heterogeneous composition. As revealed in subsequent paragraphs, the composition will vary not only because of the inherent differences of the materials from the individual barks but also owing to the procedures employed for their isolation.

Although bark and wood tissues are formed side by side in the tree, their chemical nature is unlike. Bark differs from wood by a lower cellulosic content and in the type of lignin normally present in wood. The material in barks that appears as lignin varies in amount over a wide range, whereas the lignin content of both hardwoods and softwoods varies within rather narrow limits. Bark lignin is much lower in methoxyl than is wood lignin. Ash and nitrogen contents are higher than those for wood. No constant relation exists between the contents of the other components (cellulose, pentosans, volatile acids, etc.) in wood and in bark. Wood contains no cork or lignin with carboxyl groups. A summative analysis that includes ash, extractives, holocellulose, and lignin accounts for 100 per cent of the wood analyzed. Our knowledge of the various constituents of bark tissue is not sufficiently advanced to permit a scheme of summative analysis.

TABLE 1

Bark analyses

SPECIES	MOISTURE	SOLUBILITY				PENTOSAN	FIBER	NITROGEN	ASH	REFERENCE
		Ether	Alcohol	Water	Tannin					
Alder, *Alnus viridis*		3.6	23.1	31.3	16.5		26.6	1.2	3.2	(81)
Black alder, *Alnus glutinosa* L.		5.5	22.2		8.7		33.5	1.3	3.5	(81)
Balsam, *Abies balsamae*		14.0	2.0	14.0		6.7	36.7			(25)
Basswood, *Tilia platyphyllos*		2.3	12.3	18.1	7.3		49.3	1.5	1.9	(81)
Beech, *Fagus silvatica*		3.3	10.5	14.6	3.1	28.7	36.7	1.4	8.9	(11, 12, 81)
Beech, *F. grandifolia*		1.2	13.2	4.0		13.7			8.3	(58)
						14.2				
White birch, *B. papyrifera*		2.3	12.2			16.5				
		8.0	13.6	5.8		17.1			1.5	(58)
Yellow birch, *B. lutea*		1.7	12.3	5.9		16.6			2.9	(58)
						17.5				
Dogwood, *Cornus sanguinea*		4.2	21.8	25.5	13.0		31.3	1.8	6.2	(81)
Elder, *Sambucus nigra*		1.7	6.1	15.0	2.5		36.8	1.5	5.8	(81)
Elm, *Ulmus campestris* L.		2.4	14.6	20.5			59.2	0.2	6.83	(82)
Hemlock, Western, *T. heterophylla*	9.8			13.5		5.4	33.7		1.7	(33)
Hornbeam, *Carpinus betulus*		3.4	11.4	16.9	7.9		45.0	1.7	7.1	(81)
Horse chestnut, *Pavia rubra* Lam		2.5	18.5	28.9			30.2	1.9	5.8	(83)
Oak, black	11.4			16.4		15.7	32.0		5.1	(33)
Oak, chestnut	11.9			11.1	4.6	13.4	30.4		4.3	(33, 41)
Oak, red	12.5			8.7	5.4–11.1	16.5	35.1		5.4	(33, 41)
Oak, white	12.6			10.5	2.3– 7.2	12.4	29.2		8.3	(33, 41)
Pine, European			12.3			5.7	21.5			(77)
Plane tree, *P. orientalis*		3.8	8.3	16.6	5.2		56.7	1.6	3.9	(82)
Privet, *Ligustrum vulgare*		2.6		38.3	7.0			2.3	4.4	(81)
Redwood fiber	0.0	0.2	3.1	4.6		7.7	92.6		0.3	(42)
Redwood dust	0.0	0.9	1.5	9.6		3.1				
Black spruce		2.3	4.1	10.5	6.6–12.2	8.8				(23)
		4.1	12.0	12.8		9.5			2.1	(41, 58)
Sugar maple, *A. saccharum*						9.1				
		1.9	20.3	4.6	0.4	12.6			4.3	(41)
Sycamore, *Acer pseudoplatanus*		0.9	11.5	20.7		11.4	50.5	2.0	10.5	(83)
Paper mulberry		4.8	3.9	17.3	5.5	12.9	36.4	1.0	6.7	(26)

The estimation of cellulose and lignin in bark often presents far greater difficulties than with wood, and the usual methods fail or are inadequate. Compli-

TABLE 2

Components of barks

Alder, *Alunus incana* L.	Protalnulin, p.m. 236°C.; alnulin, m.p. 260°C.; alniresinol, m.p.192–193°C.; resin acids; phlobaphenes; a sterol-like substance, m.p. 260°C.; and an unidentified substance p.m. 194–195°C. (21)
Alder, *Alnus viridis*	Protalnulin; ceryl alcohol, an alcohol $C_{33}H_{60}O_2$, m.p. 194°C., and an alcohol $C_{24}H_{42}O_2$, m.p. 250°C. (81)
Alder, black, *A. glutinosa* L.	Unidentified compounds: m.p. 83–84°C., 260°C., and 261°C. (82)
Ash, *Fraxinus excelsior*	7 per cent pectin; 20 per cent hemicelluloses (arabinose, mannose, galactose); ceryl alcohol and esters; phytosterols; resin; fats; carotenoid pigment; fraxin; mannitol; invert sugar (4, 9)
Ash, *F. silvatica*	Ceryl ester of an acid, $C_{20}H_{40}O_2$, m.p. 80°C., and a phytosterol (81)
Mountain ash, *Sorpus ancuparia* L.	Ceryl alcohol, $C_{26}H_{54}O$, m.p. 80°C.; an alcohol, $C_{35}H_{60}O$, m.p. 193°C.; an alcohol, $C_{16}H_{30}O_2$, m.p. 263°C.; resins; invert sugar; tannins; choline (15)
Aspen, sp.	Suberin (66)
Aspidosperma quirandy	Alkaloids; resins; a saponin; tannins (19)
Basswood, *P. platyphyllos*	Ceryl alcohol; sativinic acid, $C_{18}H_{36}O_6$; an alcohol, $C_{22}H_{38}O_3$, m.p. 285°C. (82)
Birch, white, *Betula alba* L.	Betulinol; tannin; gallic acid; pentosan; oil; glucoside; betuloside; betuligenol; yellow dye (73, 78)
Birch, sp. *Betulaceae*	Betulinol (75); tannins; carbohydrates; fats; waxes; alkaloids; phytosterol; fatty acids; volatile acids; pectin (60, 63, 68); suberin (12, 66)
Buckthorn	Frangularoside, $C_{21}H_{24}O_9$, m.p. 234°C., $(\alpha)_{5790} = 163°$; franguloside, $C_{21}H_{20}O_9$, m.p. 249°C., $(\alpha)_D = 134.4°$; and three glucosides containing emodol (5)
Calotropsis gigantea	Mudarol, m.p. 176–177°C.; β-amyrin, m.p. 196–197°C.; sterols; fatty acids C_{24} to C_{32}; α-calotropeol, m.p. 220–221°C.; giganteol, $C_{30}H_{50}O_2$, m.p. 224°C.; and other unidentified compounds (46)
Calycanthaceae	Fluorescent substances; a glucoside, calycanthin (43)
Chestnut, *Castanea sativa*, Mill	Ceryl alcohol; fatty acids; phlobaphenes; phytosterol, m.p. 133–135°C.; resins, invert sugar; tannins (15)
Chestnut, horse, *Pavia rubra* Lam.	A hydrocarbon; an alcohol, sitosterol; the glucoside esculetin; tannins (83)

Chuchuhuasha (So. America)	Alkaloids; a flavone pigment; a γ-pyran derivative; a catechol-type tannin; a phytosterol; vitamin C; rubber (50)
Dogwood, *Cornus sanguinea*	Phytosterol, $C_{26}H_{44}O$; platanalic acid, $C_{24}H_{40}O_3$; and a substance $C_{30}H_{52}O$, m.p. 251°C. (82)
Elder, *S. nigra*	An alcohol, $C_{23}H_{40}O_2$, m.p. 216°C.; an alcohol, $C_{27}H_{48}O$, m.p. 179°C. (81)
Elm, *Ulmus compestris* L.	$(C_{12}H_{24}O)_n$, m.p. 74–75°C.; $(C_{11}H_{20}O_2)_n$, decomposing at 240°C.; phlobaphene; phytosterol, $C_{27}H_{46}O$, m.p. 134°C. (82)
Elm, slippery, *Ulmus fulva*	Mucilage composed of galacturonic acid, *l*-rhamnose, *d*-galactose, and coloring matter (1, 22, 49)
Euonymus atropurpureus, "Wahoo"	Dulcitol; furan-β-monocarboxylic acid, $C_5H_4O_3$, m.p. 121–122°C.; euonymol, a new alcohol, $C_{21}H_{30}O_4$, m.p. 248–250°C.; tannin; coloring matter; monohydric alcohols; euonysterol, $C_{31}H_{52}O_2$, m.p. 137–138°C.; homoeuonysterol, $C_{40}H_{70}O_2$, m.p. 133–134°C.; a dihydric alcohol, atropurol, $C_{27}H_{44}(OH)_2$, m.p. 283–285°C.; citrullol, $C_{22}H_{36}O_2(OH)_2$; and palmitic, cerotic, oleic, and linoleic acids (57)
Fir, sp.	Pectin (60, 65); tannin (62); suberin (66)
Hawthorn, *Crataegus oxyacantha* L.	An alcohol, $C_{30}H_{50}O$ or $C_{31}H_{50}O$; palmitic and stearic acids; tannins; sucrose; oxalates and tartrates; and unidentified compounds $C_{26}H_{54}O$, $C_{43}H_{73}O$, and $C_{26}H_{44}O_2$ (83)
Hazel	A hydrocarbon, m.p. 16°C.; betulinol, m.p. 261°C.; lignoceryl alcohol, m.p. 76°C.; sitosterol, m.p. 135°C. (8)
Hornbeam, *Carpinus betulus* L.	Betulinol, $C_{30}H_{50}O_2$, m.p. 261°C., and phytosterol, m.p. 132°C. (7); ceryl alcohol; an alcohol, $C_{18}H_{34}O_2$, m.p. 249°C.; sativinic acid; dihydrostearic acid (82)
Larch, American	Starch; pentosans; catechol tannins; saponins; a resin; a white paraffin-like solid (37)
Locust, black, *Robina pseudacacia*	12.9–28.0 per cent protein in inner bark (30)
Lunasia *costulata* Miq.	Starch; fatty oil (stearic, palmitic, oleic, and linoleic acids, and unsaponifiable); alkaloids (16)
Mulberry, sp.	α-Amyrin, m.p. 185–195°C. (48)
Mulberry, paper, *Bronssonetia papyrifera*	Bast fibers 6.5–45.2 mm. long; resin and fatty acids; pectin; cutin; pentosan; tannin (26)
Mulberry, black, *Morus nigra* L.	Phytosterol, m.p. 132°C.; ceryl alcohol; stearic, myristic, and linolenic acids; glycerol; tannins and phlobaphenes; a sugar (21)

TABLE 2—*Continued*

Oak, sp.	Tannin (59a, 70); suberin (66)
Pine, sp.	Tannin and phlobaphenes (35, 74); pentosans, methyl pentosans, and cellulose (77); pectin (65, 68); suberin (66); fatty acid and resin acids (38); brown pigment $[C_{15}H_{13}O_5OCH_3]_x$ (39)
Australian pine, *C. calcarata* and *C. glauca*	Catechol tannins (13)
Pinus sylvestris	Lignoceric acid, $C_{24}H_{48}O_2$; palmitic, caproic, oleic, azelaic, pelargonic, and rosin acids; sitosterol, $C_{29}H_{50}O$, m.p. 137.5–138.5°C.; dihydrositosterol, m.p. 137.3–138°C.; carotene; $C_{29}H_{60}$, m.p. 59–61°C. (49).
Piscidia erythrina	Calcium oxalate (29)
Plane tree, *Platanus orientalis* L.	$C_{24}H_{42}O_2$, $C_{36}H_{74}$, $C_{24}H_{40}O_3$, and suberin (66, 82)
Privet, *Ligustrum vulgare* L.	Ceryl palmitate; ceryl alcohol; palmitic acid ester of an alcohol, $C_{25}H_{42}O$ (or $C_{26}H_{44}O$), m.p. 215°C.; phytosterol; and possible esters of behenic, arachidic, and platanolic acids (81)
Redwood, *Sequoia sempervirens*	"Bark fiber" and "bark dust"; lignin; a low-methoxyl phenolic acid; catechol tannins and phlobaphenes; a pectic acid gum; mannose, glucose, and xylose (42)
Spruce, *Picea excelsa*	Ceryl alcohol; palmiticand stearic acids; invert sugar; tannins (83)
Sycamore, *Acer pseudoplatanus* L.	An alcohol, $C_{26}H_{56}O$; phytosterol; tannins (83)
Walnut, *Juglans regina* L.	$C_{20}H_{36}O$, m.p. 63–64°C.; $C_{24}H_{42}$; $C_{22}H_{38}O$; $C_{17}H_{28}O_2$; tannin (82)
Willow, black	A glucoside (6)
Zizyphus, *Vulgaris* Lam.	Zeryl alcohol and esters; sterols; resin acids (4)

cations arise where cork and difficulty soluble extraneous materials are present or where a chemical union exists between the carbohydrates and lignin. Often the analytical results depend upon the methods used in extracting the bark prior to the actual determinations. The results of Heiduschka and Chang (26) indicate that inner bark tissues and the bark of some species may lend themselves to the application of conventional analytical procedures used for wood. Hay applied the standard wood analytical procedures to balsam bark, *Abies balsamae*, and the sum of the volatile oils, extractives, Cross and Bevan cellulose, lignin and pentosans came to 99.7 per cent (25). Undoubtedly some duplication resulted in the Cross and Bevan cellulose and the pentosan values, which would lower the total. This duplication was off-set by the loss of hemicellulose material which occurs in the Cross and Bevan cellulose determination. No reports are made of the application of the wood holocellulose procedure to barks, and where a stable union exists between the cellulose and the lignin in bark this procedure is not applicable. The chemical analyses of inner and outer bark layers, because of changes brought about by exposure, parasitic growths, and other agents, are dissimilar.

BARK LIGNIN

There is little doubt that there is present in bark a mixture of substances that appear as lignin in the conventional analysis with the concentrated mineral acid methods. These have sufficiently dissimilar properties to enable some fractionation to be made. Most bark lignins have a relatively large amount of material with a low methoxyl content, as is evident from an examination of table 3. Some lignin fractions are relatively soluble in alkaline solutions and contain significant percentages of free carboxylic and phenolic hydroxyl groups to characterize them as high-molecular-weight phenolic acids or lignin acids (42).

TABLE 3

Lignin and pentosan values for bast fibers (28)

	PENTOSAN	LIGNIN	OCH_3 IN LIGNIN
	per cent	*per cent*	*per cent*
Spruce, *Picea excelsa*	5.7	6.2	7.4
Pine, *Pinus silvestris*	5.2	12.7	1.4
Taxus baccata	16.6	12.8	0.8
Ginkgo biloba	17.1	8.5	5.6
Thuja	9.4	33.7	7.8
Eucalyptus	20.1	25.6	9.2
Red beech, *F. silvatica*	12.4	21.5	14.3
White beech, *C. betulus*	21.3	20.2	12.3
Oak, *Q. pedunculata*	19.3	12.1	9.0
Linden, *T. ulmifolia*	19.3	8.9	12.5
Maple, *A. campestre*	21.7	21.2	14.6
Poplar, *P. alba*	19.8	17.5	13.2
Hazel, *Carylus arellana*	20.3	15.6	7.7
Elder, *S. nigra*		7.7	9.4
Willow, *Salix caprea*	17.4	12.1	13.2

At least a portion of the lignin in bark is similar to wood lignin (11, 42, 77). Oxidation with alkaline nitrobenzene yields a small amount of vanillin, indicating the presence of a lignin with phenolic hydroxyl groups (11, 77). Another difference is the amount of hydroxyl groups as compared with wood lignin. Sharkov and Kalnina (67) found that pine bark lignin contained 16 per cent of hydroxyl as compared with 20 per cent for pine wood lignin.

Studies at times have failed to differentiate sharply between lignin and other bark substances, with the result that some lignin values include products formed by treatment with strong acids. In this connection, the method used in removing the extraneous materials affects the lignin value. Wacek and Schön obtained 38.5–40.3 per cent of Klason lignin from spruce when the bark was extracted only with water. The same bark gave 30.3 per cent of lignin when extracted with methanol–benzene, and 19 per cent of lignin when it was extracted successively with water, alcohol, and finally with benzene. The application of the Willstätter lignin method to this extracted bark gave a lignin yield of 20.1 per cent. The methoxyl content of the Klason lignin was 5.6 per cent and that of the Willstätter lignin 5.05 per cent (77). Sharkov (62) extracted lignin fractions from spruce and fir and concluded that they were like the products formed by subjecting the corresponding tannin to 72 per cent sulfuric acid. He also found an accumulation of lignin-like substances in the change from living bark to dead bark (61). Suberin, the major constituent of cork cells, is insoluble in cold concentrated sulfuric acid, and hence interferes with a lignin determination. Zellner (81) and Krug (33) did not include lignin determinations in their bark analyses.

Clotofski and Junge (11) obtained widely varying lignin values from birch bark, depending upon the method used; cuprammonium solution gave the lowest yield of lignin with the highest methoxyl content.

Bark lignins undergo chlorination, methylation, and acetylation reactions by conventional lignin methods. Diazomethane in ether methylates only the acidic hydroxyls, whereas dimethyl sulfate with caustic methylates all hydroxyls except those whose methyl derivatives are saponified by the caustic solution (carboxylic hydroxyls). Acetic anhydride in pyridine reacts with all phenolic and alcoholic hydroxyls. Thus it is possible to determine the percentage of different types of hydroxyl groups in a lignin and to note contrasting characteristics. This is illustrated with the alkali lignins from redwood bark and spruce wood (42):

GROUP	REDWOOD BARK LIGNIN 1	REDWOOD BARK LIGNIN 2	SPRUCEWOOD LIGNIN
Methoxyl, per cent	11.6	2.7	14.9
Carboxyl, per cent	2.0	4.4	0.4
Phenolic, per cent	4.2	7.8	3.6
Alcoholic, per cent	3.9	2.1	5.5

Redwood bark, *Sequoia sempervirens*, is separated commercially into two fractions, the "bark fiber" and the "bark dust" (42). The bark fiber is inherently a cellulose fiber. In its raw state, the cellulose appears to be combined with lignin by a linkage which is very resistant to splitting either by caustic soda or by phenol and acid. A part of the lignin appears to be closely related structurally to sprucewood lignin and also to redwood wood lignin. Two other compounds may be associated with the fiber. One is a low-methoxyl phenolic acid; the other closely resembles lignin but contains a definite percentage of carboxyl groups. The first of these is readily soluble in alkali, whereas the second is somewhat more difficultly soluble. There is a possibility that these two compounds are merely mixtures of a phenolic acid containing no methoxyl groups whatever, with redwood bark lignins in different ratios. The bark dust represents a heterogeneous mixture of broken fibers and parenchyma cells, together with a large amount of the phenolic acid found in the fiber. This particular substance probably makes up 70–80 per cent of the bark dust. Because of its behavior in the lignin analytical method, it may appear as lignin in any determination made on redwood bark, thus giving a completely erroneous picture of the lignin content of the redwood bark. The purified product recovered from the bark dust is free of methoxyl groups; it possesses both phenolic and aliphatic hydroxyl and carboxyl groups.

Varying within broad limits are the lignin, methoxyl in lignin, and pentosan contents of some bast fibers isolated by Hilpert and Knackstedt (28). Their data (table 3) show no relationship between pentosan, lignin, or methoxyl in lignin. In this connection, there is no support for the assumption that pentosans are the precursor of lignin (51).

It is evident that the various bark lignins vary in composition, possibly by closely related and isomeric substances. Bark lignin when compared with wood lignin is of increased complexity, and yet may give information that, from a fresh viewpoint and approach, can aid in solving the wood lignin problem.

DELIGNIFICATION

A great resistance to delignification of barks by the chlorination (Cross and Bevan cellulose) and alkaline and acid pulping procedures is often experienced. A commercial sulfite liquor cook of redwood bark fiber gave a pulp with 31.8 per cent of lignin (42) and a recooked soda pulp from this bark retained 28.1 per cent of lignin. Such resistance to delignification is not observed with the wood from this species. This indicates that a stable lignin–carbohydrate linkage, which is absent in wood, actually exists in bark.

A lignin–carbohydrate union in wood has been claimed by several investigators (51). Whether any lignin is actually combined with the carbohydrates in wood cannot be stated with certainty but appears unlikely. For example, it has been assumed that lignin is associated with cellulose in wood by acetal, ester, and ether linkages. Of major importance in this connection is the fact that the holocellulose contains such linkages and is isolated from wood with these linkages intact. The wood lignin molecule contains few if any carboxyl groups, whether it is isolated by acid, alkaline, or inert solvent procedures. Therefore, any ester union of lignin and carbohydrates must originate with the uronic acid fractions, the least resistant substances, and not with the more resistant cellulose fractions. Any difference in the solubility of the isolated wood cellulose and the cellulose in wood can be due to a change in colloidal state and not to a hydrolysis of chemical bonds, whereas the tenacity with which lignin is held to the cellulose in bark is indicative of an actual chemical union.

Wacek and Schön (60) obtained the best results in isolating cellulose from pine and spruce barks by the method of Kürschner and Hoffer (36). The extracted bark was treated with four successive portions of a mixture of one part itric acid and four parts alcohol. The cellulose so obtained was but faintly ellow and contained 42.2 per cent carbon, 6.4 per cent hydrogen, 0.85 per cent nethoxyl, 6.6 per cent pentosan, and 69.7 per cent alpha cellulose. By the ame procedure, Heiduschka and Chang obtained 36.4 per cent cellulose from aper mulberry bark and 54.6 per cent cellulose from the corresponding bast 26). The nitric acid method is more drastic than the chlorination method. Ience the yields of cellulose and alpha cellulose are lower than those obtained y chlorination. Both procedures give arbitrary values.

CORK

Related to the cork in bark is the suberin content. Ordinary cork from the ark of the oak tree *Quercus suber* contains 35 to 40 per cent of suberin (17). Because of its cork content the bark of this tree holds first position in commerce. The suberin content of some trees grown in Russia was determined by Sharkov nd coworkers (66). Bark from pines, fir, aspen, oak, and plane trees contained 2.0–8.3 per cent of suberin. Sequoia trees of the Socha region contained 8.1 per cent of suberin. Various kinds of birch barks are relatively high in suberin—19.7 to 38.8 per cent. The thick bark of the Douglas fir, *Pseudotsuga taxifolia,* is a rich source of cork, which can be recovered in the form of granules of fairly large dimensions by proper grinding procedures (24). Shakov holds that the transformation of living bark to the dead bark is accompanied by the development of suberin (61).

The exact composition and nature of cork is still unknown. Although cork is one of the most familiar objects in daily use by chemists, it is a substance concerning which relatively little is known chemically. Suberin, which appears to be the major constituent, is very stable chemically. It requires very drastic treatment with acids, alkalies, or oxidizing agents before it is decomposed (17). Suberin yields suberic acid (1,6-hexanedicarboxylic acid) when oxidized by nitric acid. On the other hand, Pope believes that cutose constitutes the greater part of cork. Cutose is insoluble in cold sulfuric acid of specific gravity 1.78 and in the cuprammonium solution that dissolves cellulose; it appears to be composed of stearocutic acid, $C_{28}H_{48}O_4$, with five equivalents of oleocutic acid, $C_{14}H_{20}O_4$ (53). Among the most important properties of cork is its resistance to moisture and to penetration by liquids.

EXTRACTIVES

Early investigations of bark were often directed to the study of new extractive compounds and their possible usefulness in medicinals. As a result a more extensive literature exists on the extraneous substances from barks than on other bark components. Barks of *Angostura, Cascarilla, Condurango,* and *Cinchona* have long been of interest for their production and storage of drugs. Some American medicinal barks have been described by Henkel (27).

Bark extractives contain all the types of compounds present in wood extractives. It is claimed by Richter that the chemical composition of the extractives from the outer and inner barks varies considerably, depending upon the species (58). There is greater solubility in ethyl ether than in petroleum ether. Hence the ethyl ether extracts are conveniently separated into petroleum-soluble and petroleum-insoluble fractions. Of the compounds included in table 2 those most abundant in the ether extractive are higher alcohols, fats and fatty acids, resins, hydrocarbons, and waxes. Phytosterol and ceryl alcohol are commonly present. Palmitic, stearic, oleic, and linoleic acids are the most common fatty acids found. Direct extraction of waste barks is not feasible for the production of fats and fatty acids, because the concentration is too low. The above, along with many other substances, such as coloring matters, tannins, phlobaphenes, and alkaloids, are likewise extracted with alcohol. In the aqueous extract tannins, starch, pectins, sugars, and glucosides predominate. Among the hexoses, fructose, glucose, galactose, and mannose have been reported. The pentoses include arabinose and xylose. Certain of the glucosides have fluorescent properties (43).

It must be realized that the lines of separation in the above extractives are not sharp and there is bound to be some overlapping of components. There is generally some water-soluble material in the alcohol extract. Acidulated aqueous solutions, as well as alcohol, are good solvents for the basic alkaloids.

Probably the most widely distributed compound in a genus is betulinol. This is the white pigment, "birch camphor," of birch trees. It is found in the barks of all white birches, hornbeam (7), and in other Betulaceae (63, 75). A yield of 35 per cent has been obtained from paper birch, *Betula papyrifera*, and 42.7 per cent from the white birch of Europe and Asia, *Betula alba* L. (63). Besides betulinol, the bark of *Betula alba* contains betulol, $C_{15}H_{24}O$, a sesquiterpene boiling at 255–256°C., and a hydrocarbon melting at 49–50°C. A glucoside, m.p. 190°C., is extracted with water. From sweet birch, *Betula lenta* L., is obtained oil of wintergreen (methyl salicylate) and a paraffin, $C_{30}H_{62}$ (78).

PECTIN

The uronic acid substances extracted from barks are heterogeneous compositions. Balsam, *Abies balsamae*, which is the source of the Canada balsam of commerce, is extracted with difficulty for its tannin, because of a highly gelatinous pectin-like material (23). Crude pectins of various properties are obtained from pines and firs. According to Sharkov, Girchits, and Sartaniya, pine bast contains approximately 11 per cent of pectin substances soluble in hot 96 per cent alcohol, 8 per cent of pectin substances insoluble in alcohol and soluble in water at 10–80°C., and 15 per cent of pectin substances which hydrolyze slowly in boiling water (65). Paper mulberry bark contains 9.1 per cent of pectin (26).

The pectin extracted from pine bark with ammonium oxalate solution gave upon analysis 57.5 per cent polyuronides, 12.6 per cent pentosans, 2.0 per cent methoxides, 4.2 per cent lignin, 1.6 per cent cork, and about 3 per cent volatile acids (68). A birch pectin contained 71 per cent of uronic acids and resembled the pectin from pine (68). According to Sharkov and Girchits (64), pectins from barks possess no jellying properties, although the degree of methylation may not be lower than that of pectins from beet and citrus fruit skins. Ash bark, *Fraxinus excelsior*, contains approximately 7 per cent of pectin substances together with 20 per cent of hemicelluloses (9).

A crude pectin substance containing 58.3 per cent uronic acid anhydride.

12 per cent pentosan, 11.2 per cent lignin, and 6.2 per cent ash is extracted with hot water from redwood bark. Mannan is also present as an impurity. Purification of the crude pectin with bromine water gave a pectic acid with 76 per cent uronic acid, 25.9 per cent pentosan, and 0.53 per cent methoxyl; these results indicate that it is similar to pectins isolated from other natural sources (42).

The mucilage from the bark of slippery elm, *Ulmus fulva*, is a mixture of polyuronides. Hydrolysis products are galacturonic acid, *l*-rhamnose, and *d*-galactose (1). The methyl ester of a galacturonidorhamnopyranoside has been isolated from this mucilage (22).

TANNINS

Some barks are potential sources of commercial tannin extracts, where the economic obstacles involved in collecting and transporting are not insuperable and the physicochemical properties (color, purity, astringency, degree of fixation by hides) of the extracts prepared from them are satisfactory. Among the domestic sources of tannin currently considered is the bark of western hemlock (*T. suga heterophylla*), oak, and second-growth Douglas fir. Tannin analyses of the barks of some American softwood and hardwood species were made by Levi and Manuel in 1909 (41). Their findings showed that (*1*) maple, birch, and black ash had low or negligible tannin content, (*2*) oaks, hickory, sweetgum, and yellow pine had relatively high tannin (5–11 per cent), and (*3*) white ash showed a higher tannin content than black ash. The tannin content of the barks of several eastern Canadian conifers was determined by Graham and Rose (23). Russell has made a comprehensive survey of the tannin contents of tree species of the Southeastern United States. The bark of the willow oak, *Quercus phellos*, contains over 10 per cent of tannin, and both the wood and the bark of *Castanea ashei* contain 12 per cent of tannin (59a).

The barks of Australian trees have liberal supplies of tannin unequalled in any other part of the world. Mimosa or wattle bark (*Acacia* species) has up to 40 per cent of tannin (78). Mallet bark, *Eucalyptus occidentalis*, has a tannin content of 35–52 per cent (78). The tannins from *Calletris* bark are similar to that from hemlock and their action on hide is said to be rapid and complete (72); the per cent of tannin from the different species varies from 10.7 to 30.9 (13). Another high-tannin bark is mangrove (17–48 per cent).

According to Kubelka and coworkers, the yield of tannin from pine bark can be increased by 45–75 per cent by the use of 1.35 per cent of a 1:1 mixture of sodium sulfite and bisulfite in the leach water (34). Sulfiting the tannin extract decreases the insolubles and is held to give a softer, more elastic leather. Sharkov, Tyayunova, and Kalnina obtained the best results by using 40 per cent of a 1:1 sulfite–bisulfite mixture, based on insolubles, and heating for 9 hr. at 95–98°C. (68).

In this country a vast untapped reservoir of tannin is the bark of the western hemlock, *T. suga heterophylla*. This tree is one of the principal pulpwoods in the Pacific Northwest, and its bark has long been known as a source of tanning material. Hemlock wood itself has little or no tannin. According to Pittam, who describes a pilot-plant operation (52), about 1 ton of dry bark with an average tannin content of 6.15 per cent is produced per 6 tons of pulp. Smoot and Frey obtained an average value of 15.5 per cent tannin in moisture-free bark (71). The difference is attributed to the leaching action while the logs lay in the boom. An improved process for the preparation of the tannin from hemlock bark with sodium bisulfite is described by Potter and coworkers (54).

NITROGEN COMPOUNDS

On the whole, the investigations on the nitrogen compounds in barks have not been sufficiently comprehensive to permit general conclusions.

Quinine, strychnine, and brucine are the best known of the alkaloids that occur in certain barks. The barks of certain *Strychnos* species, *S. nux vomica* and *S. ligustrina*, may contain over 7.0 per cent of strychnine and brucine (78). The early natives in South America, Guiana, and Java are said to have made an exceedingly poisonous extract from the bark of various *Strychnos* trees, which they used as an arrow poison. The poisonous properties of the extract were particularly manifested when brought in contact with a cut or scratch.

The barks of various species of *Cinchona* long were known for their antifebrile properties. These properties are chiefly due to the presence of quinine and of closely analogous alkaloids. The bark richest in alkaloids (8–12 per cent), of which 3–8 per cent is quinine, is *Cinchona ledgeriana* (10). Besides the alkaloids, the *Cinchona* barks contain fiber, starch, gum, mineral matters, coloring matters, wax, fat, traces of volatile oil, and oxalic and quinic acids (78).

Several barks of lesser known species are used in their respective countries of origin for various purposes. The barks of *Alstonia* trees occuring in India, Java, Australia, and the Philippine Islands are used as bitter tonics and for the treatment of malaria. These barks all contain alkaloid mixtures. The *Kurchi* bark of India, used as a remedy for dysentery, contains the alkaloids conessine and holarrhenine (69). The Yohimba bark of commerce contains alkaloids with physiological action resembling cocaine (69). The bark of *Aspidosperma quebracho blanco*, which grows in South America, is used as a febrifuge; it contains at least three alkaloids of similar physiological action (18). Other barks which contain more or less mixtures of alkaloids are *Aspidosperma quirandy* (19), *Lunasia costulata* Miq. (16), and the *Remijia* species (10). The absence of yohimbine and the presence of corynanthine in *P. africana* and *P. mayumbensis* support the classification of these species in the genus *Pseudocinchona* (56).

A surprisingly large quantity of a poisonous protein (12.9–27.9 per cent) is present in the bark of the common black locust tree, *Robinia pseudacacia* (30). According to Jones and Phillips, this protein was found to consist chiefly of a globulin, albumin, and a proteose. Studies on the isolated proteins found them similar in properties and composition to most vegetable and animal proteins; practically the same assortment and proportions of amino acids are present (30).

Lemesle reported that a complex protein–tannin–resin substance, which is insoluble in alcohol, water, or alkali (40), is present in the bark of *Drings winteri*.

MINERAL CONSTITUENTS

The subject of ash in barks has been reviewed by Czapek (14). The ash content and composition may vary within relatively broad limits for different trees. Often the mineral content of the bark may exceed ten times that of the wood. Manaresi and Tonegutti (45) found the ash content of the bark of the pear tree to be only about three times that of the wood.

About two-thirds to three-fourths of the ash is lime; potash may account for 10–30 per cent; sodium 0.5–2.0 per cent; magnesium 2–5 per cent; iron 0.5–3 per cent; and manganese under 1 per cent. Of the acid constituents, silica is present in largest amount; yields of 20–30 per cent of the total ash are not uncommon. A high silica content is accompanied by a low calcium content. Phosphoric acid may range from 1.5 to 4 per cent. Sulfur and chlorine usually occur only in traces. Fraymouth (20) has described a process for the recovery of oxalates from barks.

TABLE 4

Heating value of typical barks (59)

Based on oven-dry weight of bark

SPECIES	PER CENT ASH	B. T. U. PER LB.
Hardwood		9,410
Lodgepole pine		9,386
Douglas fir	0.95	10,820
Western yellow pine	1.30	9,275
Jack pine	1.51	9,870
Balsam fir	5.19	8,970
Spruce	2.53	9,590
Spruce	2.48	9,395
Hemlock	1.57	9,680

USES OF BARK

It is said that to the lumber and pulp industries bark is a necessary evil which must be disposed of in the easiest and most economical way possible. In the majority of the cases this is accomplished by burning to produce power and steam or by just burning in a waste burner. The possibility of utilizing this bark in other ways than as a fuel has been studied to a limited extent. A few special uses have been found and have been applied in isolated cases. Inasmuch as power and steam are always in demand, this was believed by Rue and Gleason to be the best use in pulp and paper mills (59). The extreme wetness of the bark (35–85 per cent) constitutes one of the chief obstacles to its utilization as a fuel.

The heating value of bark can be considered equal to or higher than that of wood. To illustrate, the heat value of Douglas fir bark, heartwood, and sapwood is respectively 10,100, 9,000, and 8,500 B.T.U. per pound of oven-dry material (79). The heating value of dry non-resinous woods is generally from 7560 to 9000 B.T.U. per pound; whereas the heat value of barks as given in table 4 normally exceeds 9000 B.T.U. The comparative fuel value of some European woods and their barks was investigated by Klein (31).

The heat value of bark and wood is affected by the extraneous materials. Mineral matters contribute little to the heating value, while rosin has about twice the heating value of wood, so that resinous woods and barks have a somewhat higher heating value depending upon the resin content.

Reduction of bark by standard chemical pulping methods ordinarily gives a low yield of pulp, accompanied by a very heavy consumption of chemicals. An exception appears to be the barks of *Crotalaria juncea* and *Sesbania sesbans.* When cooked with 20 per cent of soda they are said to give 40–44 per cent yields of pulps with high tensile strength and good alpha-cellulose content (47). Disintegration of linden bark is effected by leaching with hot lime water and shredding the product into strips (44). Spinnable fibers are claimed to be obtained from willow bark by boiling the raw bark for ½-1 hr. in a 1 per cent alkaline liquor (2). Paper mulberry bark has been used by the Chinese as a paper material for over a thousand years, and a bleached pulp comparable in quality to rag stock is described by Uchastikina and Matveer (76).

Mechanical fibers from long-fibered barks, such as redwood, are used for insulation, furniture upholstery, etc.; the bark dust is suitable for insecticide extenders. Other proposed uses for disintegrated barks are as (*a*) impregnation material for tars, asphalt, etc. (55), (*b*) wall papers, (*c*) insulating board, and (*d*) sheathing, bottle wrappers, deadening felts, chipboard, and box board (32).

Carbonization of barks gives low yields of volatile products. Comparative experiments with spruce, fir, pine, and birch gave a somewhat higher output of charcoal and tar, and less pyroligneous acid and methanol for bark than for the corresponding wood (3). Barks have practically no fertilizer value.

Long roads of research and practice must be travelled before the maximum or optimum utilization of barks can be attained. On the other hand, the complex and varied components of barks do hold possibilities. Some barks which today serve only a limited purpose will be in time explored for a much wider field of utilization.

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THE ACTION OF ELEMENTARY FLUORINE UPON ORGANIC COMPOUNDS

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I. INTRODUCTION

The element fluorine was first isolated in 1886 by Henri Moissan, Professor at the École Supérieure de Pharmacie in Paris, working in coöperation with the French Academy of Sciences. His work was characterized throughout by a high degree of scientific precision and was, incidentally, warmly commended by his colleagues (25) of that day. Fluorine is the most electronegative, and probably the most reactive, of all the known elements, so that the technical difficulculties and expense involved in preparing the gas were at first very great. As a result, the development of our knowledge of the fundamental properties, reactions, and possible uses of the element and its compounds was correspondingly slow. However, in recent years this advance has been proceeding at a rapidly accelerating rate in many directions,—inorganic, organic, industrial (as in refrigerants and insecticides), and medical (as in the study of dental caries).

In the field of organic chemistry the development has proceeded along two general lines: (*1*) indirect fluorination and (*2*) direct fluorination. The first of these consists essentially in the introduction of the element into organic compounds by means of reagents such as antimony fluorides (Swarts reaction), hydrogen fluoride in the presence of antimony or other metalloid halides as catalysts, fluoboric acid (Schiemann reaction), or in other ways. The replacement of aliphatic or otherwise reactive chlorine by fluorine, using hydrogen fluoride as the reagent, often in the presence of antimony halides as catalysts and frequently under pressure, has a number of important advantages, since the reactants are relatively inexpensive, the changes usually take place smoothly, and all of the fluorine is at least theoretically available. On the other hand, this method cannot in general be used successfully either with aromatic chlorine, or to prepare fluorocarbons or other perfluorinated compounds which are of much significance. These indirect procedures have, quite naturally, undergone important development, but it is not proposed to discuss them here.

The second method consists in the direct reaction of the gaseous element with organic compounds under a variety of conditions. This procedure leads readily to the formation of completely fluorinated organic molecules, as the result of substitution, addition, polymerization, or cleavage, which possess valuable properties and cannot in general be prepared successfully in any other way. The direct method, which is to be dealt with fully in the following pages, has certain obvious disadvantages, all of which are being greatly reduced by recent developments. Elementary fluorine is still relatively costly and its reactions are energetic and must be adequately controlled, while in substitution reactions only half of the fluorine is directly available, although the remainder may easily be recovered as hydrogen fluoride and used again.

The earlier studies dealing with direct fluorination were beset with many difficulties due to the great reactivity of the element, and it was soon found that the gas combined with most carbon compounds essentially by combustion and often explosively, yielding only carbon tetrafluoride, hydrogen fluoride, and free carbon. It required nearly half a century of slow progress before ways and means were found to control and direct these powerful reactions, but now the importance of the resulting fluorocarbons and their derivatives is increasing rapidly. Incidentally, it is beginning to look as if processes involving a combination of indirect methods with direct fluorination may have great possibilities in the preparation of such materials on a large scale. Great advances have been made along these lines during the war years, but since most of this work has not yet (1945) been published, it cannot, in spite of its interest and importance, be dealt with in these pages. However, it will be abundantly clear from what follows that the fundamental importance of elementary fluorine in its reactions with organic compounds has been firmly established.

Perhaps a short sketch covering the broad outlines of the work done in this field for the purpose of orienting the reader may not be out of place at this point. In general, direct fluorination using the element itself as the reagent can be carried out in four quite different ways, and these, listed in the approximate order of decreasing importance at the present time, are (*1*) in the vapor phase, over a metal screen packing, (*2*) in solution in an inert solvent, (*3*) in the liquid phase, by counter-current flow, and (*4*) in the solid phase, mixed with a solid, inert diluent.

The vapor-phase method, which has undergone considerable development in the writer's laboratory and elsewhere, following an earlier observation by Fredenhagen and Cadenbach, requires only a heated or cooled brass pipe reactor filled with a copper wire or screen packing, which may be silver plated. The sample, a gas or volatile liquid, which may be diluted with or carried in on nitrogen or other inert gas, is introduced into the reactor along with the fluorine, which itself may be diluted, and reacts quietly within the meshes of the packing. The products are withdrawn progressively, separated, and rectified. This method succeeds presumably because the large surface breaks up the reaction chains, and the bulk of metal rapidly conducts much of the large amount of heat produced away from the reaction zone. It is limited only by the volatility of the sample, since it is not effective in the presence of liquids. Also, if the temperature must be raised too high in order to maintain the vapor phase, molecule smashing by the energetic fluorine atoms occurs, and the product is a mixture of polymerized fragments which cannot be separated easily. Even such a mixture, however, may have valuable physical and chemical properties. Methane, ethane, acetone, benzene (to C_6F_{12}, etc.), and numerous other compounds have been successfully fluorinated by this procedure.

Direct fluorination in solution has been studied principally by Bockemüller, who used mostly a cylindrical metal reactor equipped with baffle plates and containing a dilute solution of the sample in carbon tetrachloride. Fluorine, diluted with carbon dioxide, was bubbled directly through the solutions, and in many cases reacted quite smoothly, after which the products were worked up in orthodox fashion. The reaction presumably takes place at the gas–liquid interface in the bubbles, and the procedure is limited by the relative reactivities of the sample and the solvent toward fluorine. The writer has found that such reactions may be accompanied by deep-seated chlorination of the organic molecule, and also by powerful reactions involving the solvent vapor. The butyric acids, tetrachloroethylene, crotonic acid, oleic acid, and even benzoic acid (by addition, etc.) among others were successfully reacted in this manner. Liquid fluorocarbons cannot be used in this kind of procedure, since organic compounds in general are not sufficiently soluble in them, but Freon (CF_2Cl_2) and liquid hydrogen fluoride have been used successfully as solvents by Bockemüller and Simons, respectively.

Fluorinations in the liquid phase have been the subject of a preliminary but valuable study by Miller, who employed a metal pipe U-shaped reactor either heated or cooled, in which the fluorine bubbled directly through the liquid sample, which was flowing in the opposite direction. Here, too, the reactions presumably took place at the gas–liquid interface in the bubbles, or possibly in the vapor phase. This procedure is thus far limited to compounds which are relatively inert toward fluorine, and which possess vapor pressures and melting points sufficiently low. High vapor pressures result in too powerful reactions occurring in the vapor phase. Tetra- and pentachloroethanes, as well as tri- and tetrachloroethylenes, among other compounds, have been fluorinated successfully by this procedure.

Direct fluorination in the solid phase has not often been successful, although some solid chlorocarbons like hexachlorobenzene or hexachloroethane do not react violently with the gas on contact. Ruff passed a current of fluorine through a cooled mixture of silver cyanide with an inert solid diluent, supported on a porous bed of like material, and condensed the effluent gases. Some very interesting, unusual, and highly stable compounds were isolated from the product, such as CF_3NF_2, $FCONF_2$, and the dimer $(CNF_3)_2$. This method, however, seems at present to be quite restricted in its applicability.

Turning now to the more detailed presentation, it is desirable to make clear at the outset the general plan according to which the material involved is to be organized. In the leading section there will be given a critical, specific, and fairly complete outline of the studies which have been made in the field from the time of Moissan's original discovery in 1886 to the end of 1945, so far as they have been reported in the literature. This material is to be classified chronologically by leading investigators, rather than in the more usual arrangement by methods and topics. It was felt that this would lead to a much more coherent presentation, since the number of independent contributors in this new and difficult field is quite limited, and they have attacked the problem in very different ways, often obtaining complex results involving several kinds of reactions occurring simultaneously. This means that the contributions of each scientist and his collaborators during the period covered by this review will be considered together, presumably in the order of their natural development within the group. In each case, since the material involved is relatively unfamiliar and widely scattered through the literature, it is planned to present a detailed summary of (*1*) the methods employed and (*2*) the results obtained; after which, whenever it seems practicable, the writer will endeavor to offer some appraisal of the results, either singly or collectively, in the light of recent developments and experience.

Following this more or less factual presentation, a tabular summary will be given including practically all of the well-characterized organic fluorine compounds which have been made by direct fluorination, together with the names of the investigators who first prepared them and the leading references. It is hoped that by this means a reader having in mind either a particular person, or a definite compound, will have no difficulty in locating either the discussion presented here or the original papers.

Later on, a special section is to be devoted to postulated reaction mechanisms and other similar considerations. While this theoretical presentation cannot hope to be complete or in any sense final, an endeavor will be made to give the reader some idea of how the reactions involved in direct fluorinations have most frequently been explained.

Following this, there will be given a detailed account of the construction and operation of a selected number of typical fluorine generators, since the gas has almost without exception been generated as used and has not as yet (1945) been made available commercially. A tabular summary covering all such units which have been mentioned in the literature will also be found in this section.

Finally, a description of the quantitative analysis of fluorine gas will be included, since this important and much neglected subject deserves much greater attention than has usually been paid to it.

II. THE ACTION OF ELEMENTARY FLUORINE UPON ORGANIC COMPOUNDS

A. MOISSAN,[1] Meslans, Chavanne: pioneer work in France

Quite naturally, Moissan (66) was the first to observe the effects, now so well known, of bringing elementary fluorine into direct contact with various organic substances. In the main, uncontrolled reactions, combustion, or even detonation took place, with the ultimate destruction of the organic molecule and the formation of hydrogen fluoride and fluorocarbons. Meslans (56), a pupil of Moissan's, subsequently described similar results. Moissan also studied the reactions of fluorine with the various forms of carbon and reported (64, 65) the isolation of carbon tetrafluoride by several methods, accompanied by an analysis. However, his careful efforts to purify the mixture originally obtained were actually unsuccessful, since the reported boiling point (−15°C.) bore no relation to that of pure carbon tetrafluoride (b.p. −130°C. (76)). Considerably later, Moissan and Chavanne (68) attempted to react solid methane (m.p. −184°C.) with liquid fluorine at −187°C., but even at this very low temperature the reactivity of the halogen was so great that the only result was a powerful explosion.

B. BANCROFT, Humiston, Jones, Whearty: pioneer work in the United States

On account of the many experimental difficulties involved in the original preparation of fluorine, as well as those resulting from its great chemical reactivity, no further studies dealing with the reactions of the element with organic compounds were attempted for nearly thirty years after Moissan's early work was completed. Then a program of research was begun, initiated during the first World War and conducted either under the general supervision of, or in direct collaboration with, the physical chemist W. D. Bancroft. The first results were described in 1920 by Humiston (45), who, using a Mathers generator (see below) and relatively simple equipment, brought fluorine, at times diluted with nitrogen, into direct contact with carbon and a number of simple organic compounds. He found that the carbon burned to a mixture of fluorocarbons and that acetone and paraffin hydrocarbons took fire at once, while phosgene, chloroform, and chlorocarbons were much less reactive, although free chlorine was liberated in each case. No pure products were isolated, and perhaps the most important observation recorded was that when fluorine was passed through a solution of tetrachloroethylene in carbon tetrachloride, the chlorine set free reacted with the solute to produce considerable hexachloroethane. The powerful chlorination reaction set up when fluorine is passed into solutions of organic compounds in carbon tetrachloride has since been amply demonstrated.

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[1] The name of the leading investigator in each group appears first.

Some years later, Bancroft and Jones (4), who were chiefly interested in electrochemical rather than chemical changes, reported that when benzene vapor and fluorine were mixed, there was a short induction period followed by an explosive reaction. They also noted that the fluorination of a solution of aniline hydrobromide resulted in bromination rather than fluorination, and that when aniline hydrofluoride was similarly treated, only tarry products were formed. Subsequently Bancroft and Whearty (5), and also Whearty (91), reduced the extreme reactivity of the fluorine somewhat, either by operating at lower temperatures or in solution, or by using less reactive samples. They collected considerable amounts of fluorine over an apparently little disturbed mercury surface, and then forced the gas, at a controlled rate, either directly through a liquid sample or upwards through a solid under conditions such that a liquid fluorination product might drain away and later be withdrawn from below. They found that compounds such as benzene, toluene, and anisole yielded no definite products under these conditions, and concentrated their efforts on the fluorination of solid hexachlorobenzene and a solution of 1,3,5-trichlorobenzene in carbon tetrachloride. In the first case, they obtained two products, believed to have been the compounds $C_6F_2Cl_4$ and $C_6F_3Cl_3$, respectively; in the second, they isolated one substance, thought to have been the compound $C_6HF_3Cl_2$, and their claims were supported by analytical figures. Unfortunately, however, the validity of these claims is open to question. In the first place, two of the three products boiled over wide ranges, and only one of them yielded numerically satisfactory analyses, a fact which may or may not be significant; but much more important is the admitted fact that the analytical procedure for fluorine did not give acceptable values even when applied to pure fluorobenzene. Finally, in the light of more recent work to be reviewed later, it is very doubtful if either any hydrogen or even any considerable proportion of aromatic structures could have survived the prolonged fluorination under the conditions described. Altogether it must be concluded that in all probability the products which have been described here were neither chemical individuals, nor did they consist chiefly of the compounds claimed, although the possibility that they may have contained small quantities of one or more of these compounds is not excluded. While it may appear at first glance that these researches did not yield significant results, it should nevertheless be borne in mind that they represent a definite advance into an entirely new and uncharted field, in which the expected analogies did not hold and the usual background of literature, knowledge, and experience did not exist.

C. FICHTER, Humpert, Bladergroen: fluorination compared with electrolytic oxidation

The next investigator to attempt an organic fluorination was Fichter, an inorganic electrochemist, interested in comparing the reactions of fluorine with electrolytic oxidations. In 1926 Fichter and Humpert (31) passed fluorine over solid potassium acetate, a procedure which resulted either in an explosion or a slower total destruction of the molecule, depending upon whether the salt was hydrated or not. Then they conducted the gas into aqueous solutions of potassium acetate, and obtained ethane and carbon dioxide, which would also have been produced at the anode during the electrolysis of the salt in the Kolbe synthesis of hydrocarbons. The obvious similarity was emphasized, and diacetyl peroxide was postulated as the intermediate product in the fluorination. Later, Fichter and Bladergroen (30) passed fluorine through solutions of alkali carbonates, and presented evidence indicating that percarbonates, such as $K_2C_2O_6$, were formed by a similar process. These early observations, which seem to have been made carefully, are of interest in that they shed a little light on what might be expected during the fluorination of a carboxylic acid or similar acidic compounds.

D. LEBEAU, Damiens: isolation of carbon tetrafluoride

Almost simultaneously, Lebeau, another inorganic chemist, entered the field, if but only a little way, through his study of the inert gas sometimes formed along with fluorine at a carbon anode during the electrolysis of fused metal fluorides, and also by the direct fluorination of wood charcoal. In 1926 Lebeau and Damiens (53) collected a quantity of this gas mixture, and from it isolated by repeated fractionation carbon tetrafluoride, with a molecular weight of 87.4 (calculated, 88), based on the density 3.034 at N.T.P. They described a few of the chemical reactions of this inert compound, and placed its boiling point at "close to $-150°$" (which is incorrect) but gave no analyses. Several years later Ruff (see below) repeated the work and completely characterized the compound carbon tetrafluoride; immediately thereafter Lebeau and Damiens (54) extended their work, and isolated carbon tetrafluoride (m.p. $-191°$C., b.p. $-126°$C.), but again offered no analyses. They also claimed the isolation of C_2F_6 and C_3F_8, but gave no details. It seems quite clear that while the skilled French scientists were the first to isolate carbon tetrafluoride, they cannot quite be credited with the most accurate characterization of this important compound.

E. RUFF, Keim, Bretschneider, Luchsinger, Miltschitzky, Giese, Willenberg: fluorination of carbon, graphite, carbon tetrachloride, fluoroform, and silver cyanide

Next in order to take up the study of fluorine in organic chemistry was Otto Ruff, also an inorganic chemist, who had been working with this gas for many years, and had already made important contributions to the knowledge of its reactions with many of the elements and their compounds. In 1930 Ruff and Keim (76) passed a current of fluorine over purified wood charcoal contained in a glass tube. The carbon ignited spontaneously and in time was completely consumed, leaving only a little ash. The gaseous, low-boiling products of the combustion were condensed in receivers cooled by liquid air and protected from moisture, after which the condensate was fractionated repeatedly from a small, electrically heated quartz still completely immersed in liquid air (81). There were obtained a major fraction consisting of carbon tetrafluoride; a smaller, higher-boiling portion, believed on account of its variable molecular weight and chemical inertness to be a mixture of C_2F_6 with C_3F_8; and a relatively inconsiderable residue, boiling above 0°C., whose average molecular weight approached that of C_6F_{12}. Incidentally, when the charcoal was replaced by graphite, which, on account of its molecular structure, might be expected to yield C_6F_{12} on fluorination, a greater proportion of this high-boiling residue was obtained. Also, some carbon dioxide, produced from oxygen contained in the fluorine, and a little silicon tetrafluoride from the glass were present in the mixture. Presumably owing to the relatively low fractionating power of the ingeniously constructed low-temperature still, only the major fraction was completely separated, and yielded pure carbon tetrafluoride (m.p. -186°C. ± 1°; b.p. -130°C. ± 1°) by the vapor-pressure method. Adequate analytical figures were presented, and the compound otherwise completely characterized. Metallic sodium, magnesium, and barium ignited when heated with carbon tetrafluoride, but beyond this the gas was shown to be extraordinarily inert toward a great variety of metals and reagents, as well as essentially insoluble in all ordinary solvents.

Several years later, Ruff and Bretschneider (74) continued the study of the fluorination of carbon and made the interesting observation that when either Norite or graphite was exposed to a current of the gas at a temperature somewhat below the ignition point, the carbon absorbed fluorine progressively until the mixture had the approximate composition required by the formula $(CF)_x$. Also, it was found that if the temperature of the mixture was subsequently raised, as by the onset of ignition, the product would decompose, often with explosive violence, into free carbon (soot or graphite), and a mixture of homologous carbon fluorides, possibly including C_6F_{12}. The substance was most readily prepared by conducting a current of fluorine, purified by first passing through a liquid-air trap and then through an electric discharge, for 4 hr. at a pressure of 25 mm. over a sample of carefully prepared Norite in a copper boat contained in a copper tube and maintained at a temperature of 280°C. Under these conditions the ratio of C:F in the product was 1.18:1, as determined by the increase in weight of the sample. The product was in all cases a grey powder, stable in a vacuum at room temperature, insoluble in ordinary solvents, and wet by organic liquids but not by water. It did not react with hydriodic acid, aqueous acids or bases, or even hydrogen at 400°C., but was decomposed by heat, metallic sodium, oxidizing alkali fusions, and slowly by zinc dust and acetic acid. It was considered to be a true chemical individual on account of its characteristic grey color, constancy of composition, ionic lattice structure, very low conductivity compared to that of graphite, and because it could be prepared by the action of widely varying concentrations of fluorine on both Norite and graphite, which themselves possess such widely different molecular structures. For analysis the compound was decomposed by heating to 800–900°C. for 10 min. with metallic sodium in a small iron tube welded shut at both ends, followed by determination of the fluoride ion as PbClF. The results indicated that the ratio of C:F in the compound was 1.085:1, which gave a considerable measure of support to the claim that it could best be represented by the formula $(CF)_x$. Further support was also obtained from x-ray data, which, together with the postulated mechanism of the reaction, will be discussed in Section III. Altogether, the claim that a compound of the composition $(CF)_x$ actually exists seems quite well confirmed by the accumulated mass of evidence just described, and this evidence will become practically conclusive when the experimental work involved has been repeated and checked by others.

While these studies dealing with the fluorination of carbon were in progress, Ruff and Keim (77) attempted to fluorinate carbon tetrachloride by passing fluorine through it, but found that while at room temperature little reaction occurred, at reflux temperatures dangerous explosions took place frequently, caused perhaps by the presence of chlorine fluoride in the gas. They sought to avoid this by introducing a volatile catalyst, such as iodine, arsenic, or tungsten. Since these substances react with fluorine in different valence states, they might be expected to act as intermediate fluorinating agents, or buffer substances, and so moderate the intensity of the reaction. It was found that when fluorine was introduced into the vapor of boiling carbon tetrachloride containing iodine, the mixture ignited at the tip of the delivery tube, in the presence of traces of the solid catalyst, and thereafter the fluorine burned continuously and safely in the halide vapor, with the production of a mixture of all the expected products. Afterwards, it was found even better to carry out the reaction in a slightly heated copper tube over a solid cobaltic fluoride (CoF_3) catalyst, the apparatus being so arranged that a current of carbon tetrachloride vapor, carried in on nitrogen, was mixed with the fluorine as it passed over the forward end of the catalyst layer. The products of the smooth, quiet reaction were washed with alkali, condensed, and finally fractionated repeatedly. When the molar ratio of CCl_4:F was 1:1, for example, the compounds $CFCl_3$, CF_2Cl_2, CF_3Cl, and CF_4 were formed to the extent of 32, 47, 16, and 5 mole per cent, respectively, but these proportions varied widely with the experimental conditions, as might be expected. Each of the fluorochlorides was highly purified and its physical constants reported, including analyses of the last two only, since the trichloride had already been characterized by Swarts (90), who made it by an indirect method. The compounds listed above melted at −111°C., −160°C. ± 1°, −181°C. ± 2°, −186°C. ± 1° and boiled at +24.1°C., −28°C., −80°C., and −130°C. (vapor-pressure method), respectively. The reactivities of these halides with a large number of substances were studied, and were found to increase progressively with the chlorine content.

This relatively simple and convenient process of direct fluorination over a metal halide catalyst, seemingly but little noticed at the time, actually gave real promise of significant development, as has been well demonstrated in later years. It may perhaps also be noted at this point that somewhat later in the middle of a paper describing the indirect fluorination of iodoform Ruff, Bretschneider, Luchsinger, and Miltschitzky (75) observed that the direct combination of fluorine with fluoroform (CHF_3) led to ignition and the formation of not only carbon tetrafluoride but also hexafluoroethane. This represents perhaps the first instance of the building up of a carbon chain by direct vapor-phase fluorination, a phenomenon which has since become of increasing importance.

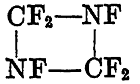
Following the studies dealing with elementary carbon, Ruff and Giese (78) fluorinated silver cyanide, and finally obtained a most interesting and extraordinarily complex reaction product. The solid sample, mixed with solid calcium fluoride to prevent violent explosions, was placed on top of a layer of pure fluorspar contained in a vertical flanged copper cylinder, maintained at 0°C. and equipped with a tight lid carrying the requisite exit tube for the products. The fluorine was led in through a copper side tube, so arranged as to deliver the gas into the underlying fluorspar layer, while the exit gases were condensed in quartz traps by liquid air and subsequently fractionated with great care. In this way there were obtained, along with carbon tetrafluoride, silicon tetrafluoride, carbonyl fluoride, carbon dioxide, nitrous oxide, nitric oxide, and polymerized residue as by-products, essentially three main fractions. The first was a colorless mixture (m.p. about −130°C., b.p. about −78°C.), about 28 per cent by volume; the second, a colorless eutectic (m.p. −128°C., b.p. −40°C.) with a composition and molecular weight corresponding to the formula $(CNF_3)_2$, about 60 per cent by volume; and the third, a dark blue substance (m.p. about −150°C., b.p. about −80°C.) with constants corresponding to the formula CF_3NO, about 2 per cent by volume. It was shown that the oxygen in this last compound came most probably from silver nitrate in the original cyanide. The first mixture could be resolved neither by fractionation nor by any ordinary chemical means, since it was completely inert toward all kinds of laboratory reagents except heated sodium. However, when this material was fluorinated at red heat, it yielded pure carbon tetrafluoride and nitrogen trifluoride in approximately equal proportions, together with pure hexafluoroethane; on this basis it was concluded that the original mixture consisted solely of the compounds CF_3NF_2 and C_2F_6.

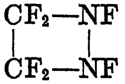
In a subsequent paper Ruff and Giese (79) studied very carefully the second and also the main fluorination product of silver cyanide, referred to above as a eutectic (m.p. −128°C., b.p. −40°C.) having the exact composition corresponding to the formula $(CNF_3)_2$. This material was presumably formed by the dimerization of $FN{=}CF_2$, which, in turn, might be expected to result from the action of fluorine on $AgN{=}C$, and was recognized as a mixture of isomers since it would react in part, but not completely, with certain reagents. It may have six possible structures, as follows:

$CF_3N{=}NCF_3$
I

$CF_3N{=}CFNF_2$
II

$CF_2{=}NCF_2NF_2$
III


IV


V

$NF_2{-}CF{=}CF{-}NF_2$
VI

of which I, II, and VI can exist in *cis* and *trans* forms. First a sample of the gas was treated with liquid ammonia in a sealed tube, which caused the appearance of fluoroform and nitrogen in the gas phase, presumably owing to the ammonolysis of II; and also of more fluorine in the residue than could be accounted for in this way, a result which was attributed to the simultaneous destruction of III and V by the reagent. The unchanged gas which had survived this treatment was then shaken with metallic mercury, resulting in the very slow appearance in the gas phase of a compound containing two fluorine atoms per mole less than the original, which was considered due to the reaction of the mercury with IV. After this, a new sample was treated first with mercury, which caused a two-stage reaction, fast at the outset and then slow, which was thought to be due to the successive attack of the metal on V and then on IV. Finally, the surviving gas was treated with liquid ammonia, with the formation of fluoroform and nitrogen as before. These results were well supported by exhaustive analyses, and it was postulated that the ammonia reacted with carbon-nitrogen double bonds, while the mercury attacked the $>$NF linkage exclusively. Ultimately it was concluded that the original mixture contained at least five of the six possible isomers; by means of a complex calculation frankly described as "more or less" convincing, it was estimated that isomers I, II, III, and IV (or V) were present in the mixture to the extent of about 52, 15, 26, and 8 per cent, respectively.

It should be noted here that considerably later on Ruff and Willenberg (82) reported the synthesis of $CF_3N{=}NCF_3$ (I) by the action of iodine pentafluoride on iodine cyanide, and characterized it as a slightly greenish, very inert gas (m.p. −133°C., b.p. −31.6°C.). Some $(CF_3)_2NH$ (m.p. −130°C., b.p. −6.2°C.) was also formed in this reaction.

In a final article of this series, Ruff and Giese (80) carried out an exhaustive study of the third portion, described above as a blue compound (m.p. about −150°C., b.p. about −80°C.) having the exact composition corresponding to the formula CF_3NO. This material, after very careful purification, was finally found to melt at −196°C. (freezing-point curve), and to boil at −84°C. (vapor-pressure curve). The very low melting point, together with the wide liquid range, gave rise to the suspicion that this portion, too, might be a mixture of the two possible isomers CF_3NO (blue) and $FCONF_2$ (colorless), and it was therefore treated with a wide variety of reagents in order to effect a separation if possible. It was found that zinc dust and acetic acid, platinum sponge and hydrogen, and manganese heptoxide in acetic acid each destroyed approximately half the sample, leaving the colorless isomer $FCONF_2$, while dilute sodium hydroxide and adsorption on charcoal completely rearranged the blue compound into the same colorless isomer, without change of volume or molecular weight. The completely pure $FCONF_2$ melted at -152.5°C., boiled at -82.4°C., and was so extraordinarily stable that it would react neither with any ordinary reagent tried, nor even with elementary fluorine itself at red heat, but only with the heated metallic sodium used in the analysis. It was recognized that such a compound, when regarded as a derivative of carbonyl fluoride, would certainly not be expected to possess any such stability, especially toward hydrolysis, but no further explanation of the phenomenon was given. From the corresponding vapor-pressure curves, it appeared likely that the boiling point of pure CF_3NO should be very close to that observed for the blue mixture, or -84°C., but the melting point should be much higher, and probably close to -150°C.

Summing up the results of these last three papers, it appears that Ruff and Giese have convincingly demonstrated the formation of the new and interesting compounds CF_3NF_2, $(CNF_3)_2$ (several isomers), and $FCONF_2$ by the direct fluorination of silver cyanide, and have given a considerable description of the physical and chemical properties of these substances. However, only one of them, $FCONF_2$, could be obtained entirely pure and free from other molecular species.

It remains to be said here that the extensive work of Ruff and his collaborators, of which the above sketch dealing with carbon compounds represents but a small fraction, is in general characterized by a measure of insight, experimental skill, and painstaking care which confer upon it a high degree of scientific trustworthiness.

F. ISHIKAWA; Murooka

The next year after Ruff began to study the fluorination of carbon and its compounds, Ishikawa and Murooka (46) reported the approximate fluorine concentrations, varying from 8–16 per cent, presumably in air, required to ignite a number of common organic liquids. They did no further work with elementary fluorine, but later on became interested in oxygen fluoride.

G. BOCKEMÜLLER: fluorination of n- and isobutyric acids, cyclohexane, n-hexadecane, tetrachloroethylene, crotonic acid, oleic acid, and benzoic acid (all in solution)

It was in 1933 that the organic chemist Wilhelm Bockemüller (10), after a considerable experience with indirect fluorination using lead tetrafluoride (29) and phenyl iodide fluoride ($C_6H_5IF_2$) (9) as reagents, published the results of an intensive study of the direct fluorination of organic compounds in carbon tetrachloride and dichlorodifluoromethane solutions. He used a well-designed closed generator (see table 2 below), and when carbon tetrachloride was the solvent, a fluorination vessel of special design, made entirely of pure nickel (see figure 1). It consisted of a cylindrical container, equipped with a well-fitting lid, which carried both a short exit tube and a long central inlet lead, to which were attached several parallel baffle plates, as shown. These were so arranged that the incoming fluorine, rising from below, would be caught under each plate for more or less time, and so maintained in contact with the solution as long as possible, since the gas was but slightly soluble in it. On the other

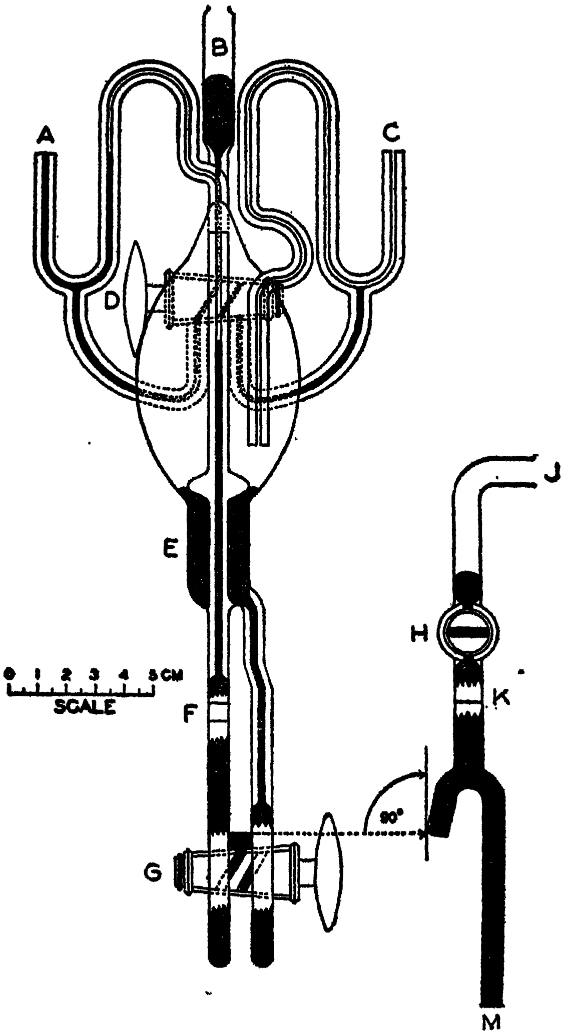


Fig. 1. Apparatus for the direct fluorination of organic compounds in carbon tetrachloride solution (Bockemüller).

hand, when dichlorodifluoromethane was used as solvent, the reactions were carried out in an ordinary glass flask, cooled to about −80°C., either from the outside, or by solid carbon dioxide contained in the solution, in which case the flask was suitably insulated. The solvent was condensed into the flask from a cylinder, and the gas led in through a nickel tube.

The fluorine was supplied at a current efficiency of 70–85 per cent (calculated, without this factor, 0.1 mole fluorine, or 3.8 g. = 5.36 amp.-hr., using a current of 4–5 amp.), and was diluted at the generator with approximately an equal volume of inert gas, usually carbon dioxide. The samples were often used in 0.1 mole quantities, and the fluorine, which was not freed from entrained hydrogen fluoride, was usually used in amounts greater than could be absorbed by the respective solutions, and always introduced in as small, finely divided bubbles as possible. Also, the nearly closed system used was provided with adequate safety devices, including an automatic alarm to warn against rising pressure due to obstruction of the lines, and was so arranged as to prevent, to a great extent at least, the access of air and moisture during the fluorinations. Finally, attention was called to the fact that while the solvents appeared to be practically inert to fluorine at the temperatures used, nevertheless a mixture of fluorine and carbon tetrachloride vapor could detonate with great force without warning or apparent cause.

The substances fluorinated may be divided into three classes,—namely, saturated, unsaturated, and aromatic; the more important results obtained will be described in that order. First, an approximately 2 per cent solution by weight of *n*-butyric acid (0.1 mole) in carbon tetrachloride was fluorinated in the nickel apparatus at 0°C. for 10 amp.-hr. The solution was refluxed to remove hydrogen fluoride, and extracted with sodium bicarbonate solution; then the aqueous layer was acidified and continuously extracted with ether. Next, the solvent was dried and removed, leaving a yellow oil, which was fractionated under reduced pressure. In this way there were obtained two significant portions, one a colorless liquid (b.p. 80–98°C. at 13 mm.), and the other a small brown oily residue. The main fraction had the analytical composition required by monofluorobutyric acid, and was obtained in about 30 per cent yield, but it was clearly a mixture, in which the halogen might occupy the α-, β-, or γ-position, although it yielded a *p*-phenylphenacyl ester which melted sharply at 92°C. after many crystallizations. This portion was therefore refluxed with sodium hydroxide, and the mixture subsequently acidified, yielding a brown oil and an aqueous solution. The oil, after extraction, rectification, and ultimate crystallization, finally yielded pure crotonic acid, $CH_3CH{=}CHCOOH$ (m.p. 72°C.). The solution, believed to contain butyrolactone, was oxidized with bichromate and sulfuric acid, and finally yielded, after suitable purification, pure succinic acid, $HOCOCH_2CH_2COOH$ (m.p. 183°C.). These findings, respectively, furnished quite convincing evidence that the original monofluorobutyric acid consisted of a mixture of the β- and γ-isomers. Finally, pure α-fluorobutyric acid was synthesized by the indirect fluorination of the corresponding α-bromo ester with a mixture of silver fluoride and calcium fluoride at about 150°C. This acid yielded a pure *p*-phenylphenacyl ester, which, although it melted at 91.5°C., was not identical with the corresponding derivative obtained before; this result makes it probable, but not entirely certain, that the monofluorobutyric acid mixture did not contain any of the α-isomer.

The small oily residue from the distillation of the original crude product was considered of interest, since it finally solidified and on purification yielded pure succinic acid (m.p. 184°C.). From this it was concluded that the fluorination must have produced some of the compound $CF_3CH_2CH_2COOH$, which, on working up, was hydrolyzed to succinic acid. This is extremely unlikely however, on account of the great stability of the resonating —CF_3 group which is now well established. It would seem probable that this succinic acid might have been formed from the oxidation of butyrolactone or of the corresponding hydroxy acid by peroxides produced by the action of the fluorine on the carboxyl group of the original butyric acid.

n-Butyryl chloride, and the mixed *n*-butyric acetic anhydride, when fluorinated under approximately the same conditions gave essentially the same results as the free acid; but in contrast, isobutyric acid yielded a crude fraction (b.p. 76–84°C. at 13 mm.) which, on further distillation, furnished analytically pure β-fluoroisobutyric acid (b.p. 80–82°C. at 13 mm.). Since this compound could readily be brominated by the Hell–Volhard–Zelinsky method, in which only the α-position of a carboxylic acid is attacked, it was clear that in this case the β-position was the only one remaining for the fluorine atom.

From work with butyric acid and its derivatives, Bockemüller quite properly concluded that the mechanism of fluorination, which led only to β- and γ-substitution, must be fundamentally different from that of chlorination or bromination, which in like circumstances leads mostly to α-substitution.

Two other saturated aliphatic compounds studied were cyclohexane and *n*-hexadecane, $C_{16}H_{34}$ (m.p. 15–16°C., b.p. 150–155°C. at 13 mm.). The former (0.12 mole) was fluorinated in dichlorodifluoromethane solution (7 per cent) in the glass apparatus at −80°C. for 10 amp.-hr. After the evaporation of the solvent the residual oil was steam distilled, fractionated, and finally yielded analytically pure fluorocyclohexane, $C_6H_{11}F$ (b.p. 100°C.) in about 33 per cent yield. This compound was described as being very unreactive. The second hydrocarbon (0.044 mole) was halogenated in carbon tetrachloride solution (2 per cent) in the nickel vessel at 15°C. for 10 amp.-hr. Then the mixture was washed with sodium bicarbonate solution, dried, and the solvent removed, leaving a yellow oil which was fractionated at reduced pressure. The final product (m.p. about 7°C., b.p. 104–112°C. at 0.2 mm.) had an analytical composition close to that required by the monofluoro derivative, $C_{16}H_{33}F$, and undoubtedly consisted of a mixture of isomers.

Turning now to the unsaturated compounds studied, an approximately 17 per cent solution of tetrachloroethylene (0.18 mole) in dichlorodifluoromethane was fluorinated at −80°C. in the glass apparatus with continuous stirring for 16 amp.-hr. The yellow oil remaining after the evaporation of the solvent was first refluxed to remove unstable and volatile impurities, and then washed with alkali, dried, and fractionated, in part at reduced pressure. In this way three significant portions were obtained which boiled at 87–93°C., 30–50°C. (13 mm.), and 139–142°C. (13 mm.). The first of these, which mostly solidified at room temperature, was the chief product of the reaction, and after further fractionation yielded analytically pure $CFCl_2CFCl_2$ (m.p. 26.5°C., b.p. 91°C.), which was the expected product. The second portion, present only in small amount, also largely solidified on cooling, and after recrystallization melted at about 100°C. This sample on sublimation yielded analytically pure CCl_3CFCl_2. Both of these were known compounds, having been originally synthesized by Booth, Mong, and Burchfield (14), and had a characteristic odor similar to that of camphor. The former substance has been obtained in an exceptionally high state of purity (m.p. 28.1°C., b.p. 91.50°C. ± 0.05° by Horvorka and Geiger (44)). The third fraction, a colorless oil which was formed in considerable amount, consisted of the analytically pure dimer, $C_4F_2Cl_8$ (b.p. 139–142°C. at 13 mm.), which, under the circumstances, was in all probability 1,4-difluoroöctachlorobutane, $CFCl_2$-$CCl_2CCl_2CFCl_2$, although this structure was not proved by direct chemical evidence.

Next, a practically 5 per cent solution of crotonic acid ($CH_3CH{=}CHCOOH$) (0.23 mole) in carbon tetrachloride was fluorinated at 0°C. in the nickel apparatus for 10 amp.-hr. The gas was completely absorbed in this case, and a moderate evolution of hydrogen fluoride took place. After removal of the solvent, the crude product was distilled at reduced pressure, and yielded a colorless distillate (b.p. 80–115°C. at 13 mm.), which mostly solidified, and a large brown oily residue.

The former product, which contained large amounts of unreacted crotonic acid, was strongly oxidized with concentrated nitric acid, and the product, after an ether extraction of its solution in alkali to remove neutral impurities, was finally recovered from an acid solution by a second ether extraction. The partially purified acid was then distilled at reduced pressure and yielded crude α,β-difluorobutyric acid (b.p. 90–115°C. at 13 mm.) in about 14 per cent yield. This material was now refluxed with pyridine, again recovered, and refractionated at reduced pressure. This time two significant portions were obtained, which boiled at 90–105°C. (mostly 100–102°C.) and 105–110°C. at 13 mm. The first of these, after final purification by repeated oxidation with nitric acid until it was stable to permanganate, yielded an analytically pure, colorless liquid, $CH_3CHFCHFCOOH$, which would not solidify at 0°C., but was still contaminated by small amounts of its solid stereoisomer. The second fraction, which solidified on cooling, finally yielded after several crystallizations a relatively small amount of analytically pure, white, crystalline $CH_3CHFCHFCOOH$ (m.p. 81°C.), which was also unaffected by alkaline permanganate solutions. The structure of this compound was further substantiated by refluxing it with aqueous sodium hydroxide, which readily converted it into analytically pure, white, crystalline α-fluorocrotonic acid, $CH_3CH{=}CFCOOH$ (m.p. 112°C.), as would be expected.

The large oily residue from the original distillation of the fluorinated crotonic acid was shown to consist largely of a dimeric difluorodicarboxylic acid, presumably having the composition required by the formula $C_8H_{12}O_4F_2$. It was esterified with methyl alcohol, but since the ester could not be completely purified by distillation, it was saponified with alcoholic alkali, which also removed all the fluorine, leaving an unsaturated dimeric acid. This product, after crystallization, melted at 240–244°C. (with decomposition) and gave analytical figures corresponding to the formula $C_8H_{10}O_4$, but the exact structure was not determined.

Oleic acid (0.1 mole) was also fluorinated in 7 per cent solution in carbon tetrachloride at 15°C. in the nickel vessel for 10 amp.-hr., all the gas being absorbed and some hydrogen fluoride evolved during the process. After the removal of the solvent, the oily residue partially solidified, and, on filtration in the cold, yielded a small quantity of crude crystals; these crystals, on repeated recrystallization from ligroin (70°C.), finally yielded analytically pure 9,10-difluorostearic acid (m.p. 81°C.), which was stable to alkaline permanganate solution. In one run, repeated fractional crystallization yielded a small quantity of an analytically pure stereoisomer of this compound, melting at 84–85°C. and not identical with the compound melting at 81°C.

Turning next to the fluorination of aromatic substances, Bockemüller was confronted with a more difficult problem. Benzene, bromobenzene, *p*-dibromobenzene, *m*-dinitrobenzene, quinone, 1,3,5-trinitrobenzene, and 2,4,6-trinitrotoluene gave no definite results, while benzoic acid, fluorinated in the nickel apparatus, yielded only an amorphous, polymeric solid mixture of variable composition, containing about 30 per cent of fluorine and decomposed even by carbonate solution, together with lachrymatory oils, resin, and unchanged sample. Since these phenomena appeared to be caused by polymerization, it was decided to inhibit this process by saturating any reactive intermediates present with bromine. Accordingly, an approximately 2.5 per cent solution of benzoic acid (0.08 mole) was treated with 0.12 mole of bromine, and fluorinated in the nickel apparatus at 0°C. for 15 amp.-hr., with the evolution of hydrogen fluoride. After removal of the bromine, the mixture was extracted with bicarbonate solution, the solvent removed under reduced pressure, and the oily residue steam-distilled, yielding a yellow, pungent, lachrymatory oil which was then repeatedly fractionated. In this way two significant liquid fractions were obtained, which boiled at 110–120°C. and 150–160°C. at 13 mm. and had analytical compositions not far from those required by the formulas $C_6H_6F_4Br_2$ and $C_6H_6F_3Br_3$, respectively. Relatively little polymerization occurred under these conditions, and the products seemed to be halogen derivatives of cyclohexane, formed by the direct addition of both fluorine and bromine to the aromatic nucleus, accompanied by the elimination of the carboxyl group. Accordingly, an attempt was made to aromatize some of the material by heating it with a zinc–copper couple. No definite results were obtained, except that the final product seemed to undergo nitration in preference to oxidation when treated with a nitrating mixture. This work constitutes the first fairly definite evidence that the primary attack of elementary fluorine on the aromatic nucleus is one of addition, rather than substitution.

The investigation which has just been outlined, which seems to have been summarily cut off by the events leading up to the recent war, represents not only an extensive, but also a markedly significant contribution to the field, and undoubtedly constitutes the most definite advance made up to this time by an organic chemist. It is characterized throughout by the solid analytical support and careful attention to detail which create confidence that the results which have been reported are substantially correct. It is to be regretted, however, that the investigator apparently did not have at his disposal certain types of equipment, such as modern highly efficient fractionating columns, by means of which the results obtained with difficult mixtures could have been considerably extended and improved. The more theoretical aspects of Bockemüller's work will be discussed further in Section III.

H. BIGELOW, Pearson, Cook, Miller, Calfee, Fukuhara, Young, Hadley: fluorination of hexachlorobenzene, hexachloroethane, tetrachloroethylene, ethane, methane, ethyl chloride, acetone, and benzene (mostly in the vapor phase)

The present writer and his students entered the field of organic fluorinations in 1933, in effect simultaneously with Bockemüller. In an introductory paper Bigelow, Pearson, Cook, and Miller (7) described the action of fluorine as produced by a Mathers type of generator on a number of aromatic compounds, such as toluene, naphthalene and its derivatives, and benzil, among others. They used copper apparatus for the most part, and carried out the reactions (*1*) in carbon tetrachloride solution or in suspension at 0°C. under ordinary conditions, and (*2*) with a pure liquid or a carbon tetrachloride suspension at 0–25°C., in the absence of any appreciable vapor phase. Under such conditions the fluorinations were readily controlled and proceeded in a smooth and regular manner. No new compounds were isolated in this study, but a number of significant observations were recorded, as follows: First, the fluorine liberated chlorine from carbon tetrachloride even at 0°C., with the formation of a chlorinating mixture sufficiently powerful to convert toluene ultimately into hexachlorobenzene in considerable yield. Second, the fluorine attacked the aromatic nucleus by an essentially different mechanism from chlorine, often yielding polymerized products which were sensitive to very dilute alkali but extremely resistant to acid oxidizing agents, even hot fuming nitric acid. Third, the fluorine could be passed directly through a liquid hydrocarbon without violent reaction, provided there was no appreciable vapor phase present; nevertheless partial combustion occurred, yielding only hydrogen fluoride and free carbon. Also, the course of the various reactions was followed by numerous analyses, which indicated that the solid polymeric products obtained from naphthalene, benzil, and their derivatives contained roughly 25–60 per cent of fluorine. Finally, it was noted that neither simple nor substituted phthalic or benzoic acids were obtained from these products on strong oxidation, indicating not only that both rings were always involved, but also that the process could hardly have been ordinary substitution. Subsequently much of this preliminary work has been well substantiated, and the fact that the primary attack of fluorine on the aromatic nucleus consists not of substitution but presumably of addition, as postulated by Bockemüller (see above) at the same time, has become practically certain.

Bigelow and Pearson (6) then undertook the fluorination of hexachlorobenzene, C_6Cl_6, a symmetrical, negatively substituted, relatively unreactive aromatic compound containing no hydrogen, which seemed well suited to the purpose. At first, an approximately 6 per cent suspension of the sample (0.18 mole) in carbon tetrachloride was fluorinated in a glass apparatus at 0–20°C. for 105 amp.-hr., after which the clear solution was washed, dried, and the solvent removed, leaving a crude, highly complex liquid product in about 60 per cent yield by weight. This material was accumulated, distilled, and a portion finally fluorinated directly, without solvent, in a metal outfit, for a total of 133 amp.-hr. more. During the process, which took place smoothly and was designed to carry the reaction to completion as far as possible, the sample was subjected to further rectifications, and two significant fractions, one weighing 30 g. (b.p. 100–115°C. at 7 mm.) and one weighing 62 g. (b.p. 135–160°C. at 1 mm.), were collected. When small samples of these oils were emulsified with water and allowed to stand in the cold, solids separated, which, after further purification from dilute alcohol, melted at 94–96°C. and 113–114°C., respectively, and had analytical compositions corresponding closely to those required by the formulas $C_6F_4Cl_6$ and $C_6F_6Cl_6$. This was considered to be the first definite evidence in favor of the addition hypothesis. In comment it is only fair to say that while this evidence appears legitimate, the analytical samples represented only a minute proportion of the materials involved, and the part played by the water in the solidification has not yet been determined with certainty. Also, the evidence presented by Bockemüller (see above) in support of the same postulate, even though he did, to be sure, use fluorine and bromine together, is at least equally convincing.

By this time several important factors involved in these reactions had been recognized: namely, that fluorine produced by a Mathers-type generator was often highly contaminated with other gases; that this halogen reacted with almost any organic solvent, frequently including carbon tetrachloride even at relatively low temperatures if in the presence of an organic solute; and finally, that all direct fluorinations seemed to yield highly complex reaction products. In fact, it appeared probable that during the course of a fluorination in carbon tetrachloride there might well be present, in addition to the fluorine, the solvent, the solute, and the reaction products, such substances as oxygen, oxygen fluoride, chlorine, chlorine fluoride, and hydrogen fluoride as well as others in lesser amounts. It was hardly surprising that the earlier results had been baffling and often disappointing. Accordingly, Miller and Bigelow (60) (see below) designed and placed in service a new closed generator, shown to deliver fluorine (94–99 per cent pure by analysis) at an average electrical efficiency not far from 80 per cent. Also, it was decided to fluorinate in the vapor phase, thus avoiding the use of solvents altogether, and to study simpler molecules, in the hope of obtaining less complex products.

On this basis, Miller, Calfee, and Bigelow (61) carried out the vapor-phase fluorination of hexachloroethane (C_2Cl_6) over a copper-gauze catalyst, by a modification of the procedure originally described by Fredenhagen and Cadenbach (33) (see below). It was expected that the heat-conducting metal would moderate the reaction by reducing local overheating, and that the presence of a surface layer of cupric fluoride might promote the change. The simple apparatus consisted of a glass tube heated by sliding air baths. Fluorine from the closed generator, at the rate of about 1.5 liters per hour and diluted with nitrogen in the molar ratio of 1:1.6, entered the reactor at one end through the meshes of a tightly wound copper-gauze roll. Beyond this came the solid sample, which was followed by a second similar roll, and the products passed out of the tube at the other end, where they were condensed by a freezing mixture. In operation, the second gauze was heated, then the first, and finally the sample was progressively sublimed into the fluorine, where it reacted smoothly and steadily. The crude liquid product, condensed at the rate of 1.6 g. per hour, or 75 per cent by weight of the hexachloroethane consumed, was washed, dried, rectified in a Podbielniak still, and finally gave a 20 per cent yield of pure *s*-difluorotetrachloroethane, $CFCl_2CFCl_2$ (m.p. 24–25°C., b.p. 92°C.). This known compound had the correct analytical composition and physical constants, making its identity certain. This reaction was the first case to be described in which the vapor-phase fluorination of an organic compound containing more than one carbon atom yielded a pure, individual product.

In addition, it was found that when the vapor of tetrachloroethylene (C_2Cl_4) was carried into the same apparatus on a current of nitrogen, and fluorinated over the copper-gauze catalyst, the same compound was formed, as would be expected, the yield being also 20 per cent of the calculated amount.

Since the vapor-phase fluorination of a simple molecule such as hexachloroethane over a metal catalyst had proved both convenient and relatively successful, the next logical step was to apply the procedure to a hydrocarbon, in spite of our earlier conviction that this would lead to violent and uncontrollable reactions. Consequently, Calfee and Bigelow (20) permitted fluorine and ethane to mingle within the meshes of a copper wire-gauze packing, contained in a brass pipe, and so arranged that the halogen entered behind the hydrocarbon in respect to the direction of the gas flow. This very simple device proved surprisingly successful, since the reaction took place smoothly, quietly, and continuously, so long as the reactants were supplied; and neither the relative proportions of the reacting gases, nor their order of mixing, nor even the nature of the sample had any appreciable effect on the general result. In operation, fluorine from the closed generator, at the rate of nearly 2 liters per hour, was mixed with ethane in the approximate molar ratio of halogen to hydrocarbon as 3:2. The gaseous products then passed over sodium fluoride contained in two similar brass pipes in series, to absorb quantitatively all the hydrogen fluoride which had been formed, after which they were condensed by liquid air in a trap carefully protected from air and moisture. The fluorinated ethane, which was formed at the rate of about 1 cc. per hour, was accumulated for 18 hr., passed through alkali, dried, and carefully rectified, using a Podbielniak low-temperature distilling column and a simplified form of the excellent fractionating apparatus described by Booth and Bozarth (13). Finally, three definitely significant fractions were obtained: 4–5 cc. (b.p. −125°C.); 1 cc. (b.p. −88°C.);

and 5–6 cc. (b.p. −78°C.). These fractions were undoubtedly carbon tetrafluoride, unchanged hexachloroethane, and hexafluoroethane, respectively. Two other portions were also described (b.p. −84°C. and −50° to −48°C.), which were believed to be CHF_3 and either CH_3CF_3 or CH_2F_2, respectively. However, the presence of these compounds in the fluorination product has not been confirmed by subsequent work, and it is quite likely that the second portion actually consisted of the then unknown compound CF_3CHF_2, which boils at −48.5°C. The procedure described here gave great promise of general application in the study of organic fluorinations, a promise which has since been amply fulfilled. It also represents the first instance in which a hydrocarbon was acted upon by elementary fluorine to produce a definite, known chemical individual other than carbon tetrafluoride.

Continuing the study of the fluorination of ethane, Calfee, Fukuhara, and Bigelow (21) carried out the reaction using progressively varying proportions of the reacting gases. The apparatus and operation were much the same as before, except that the respective volumes of the reactants were carefully measured or estimated, the catalyst packing was changed to small wire-gauze squares, and the crude condensate, which in most cases contained no free fluorine, was fractionated directly, using the completed Booth–Podbielniak low-temperature fractionating unit. This last consisted of a Podbielniak column equipped with liquid-air cooling, thermocouple, and automatic recorder, a flexible condensing system, melting-point apparatus, and gas-density balance for determining molecular weights during the distillations.

It was found that as the molar ratio of $F:C_2H_6$ varied from 1:1 to 6:1, the relative amounts of carbon tetrafluoride and hexafluoroethane produced increased, and the quantity of recovered ethane decreased as would be expected, but the proportion of partially fluorinated ethanes formed boiling higher than hexafluoroethane (−78°C.) never rose above 30 liquid volume per cent of the total condensate, and was usually far less than this value. Clearly, the reaction would have to be further moderated in intensity before these latter compounds could be isolated and studied. However, two interesting azeotropic mixtures (b.p. −92°C. and −104°C.) were isolated at this time. They were shown to consist of hexafluoroethane and ethane in the approximate molar ratio of 1:2, and of hexafluoroethane, ethane, and silicon tetrafluoride in the ratio of 1:6:1, respectively. The silicon tetrafluoride must have resulted from the action of some unabsorbed hydrogen fluoride on the glass apparatus.

As a result of the preceding work, Young, Fukuhara, and Bigelow (92) undertook a study of the influence of progressive nitrogen dilution on the fluorination of ethane, in order to moderate the reaction and to secure greater amounts of the partially fluorinated ethanes as products. The apparatus and procedure remained unchanged, except that measured amounts of diluent nitrogen were introduced into the fluorine line. The crude liquid-air condensate, usually obtained at the rate of 1.5–2.3 cc. per hour and well protected from air and moisture, was rectified in the Booth–Podbielniak unit. After many runs at molar fluorination ratios ($F_2:C_2H_6$) of 2:1 and 4:1, and dilution ratios ($F_2:N_2$) ranging from 1:1 to 1:20, it was found that when the former was 4:1 and the latter 1:3, corresponding to a reaction ratio ($F_2:C_2H_6:N_2$) of 4:1:12, the proportion of partially fluorinated ethanes formed rose to 90 liquid volume per cent of the total condensate. Under these conditions, the fluorination product contained no hexafluoroethane, no unchanged ethane, and but little carbon tetrafluoride. The material boiling above −78°C. was then accumulated, and repeatedly rectified with great care and at various pressures. Finally, pure CF_3CHF_2 (m.p. −103°C., b.p. −48.5°C.), CHF_2CHF_2 (b.p. −23.5° to −22.5°C.), and CHF_2-CH_2F (m.p. −84°C., b.p. −3°C.) were isolated from the complex mixture. Of these all gave correct analytical figures and molecular weights, but only the first was a new compound, the others having been made earlier by indirect methods. These results showed definitely that the vapor-phase fluorination of ethane under mild conditions yields, in addition to carbon tetrafluoride, hexa-, penta-, tetra-, and trifluoroethanes, but it is also of interest that none of the corresponding mono- or disubstitution products could be detected in the mixture. It is known, however, that the compounds CH_3CH_2F and CH_2FCH_2F are to some extent thermally unstable, which may account for the fact that they were not formed, but this is speculative and would hardly account for the fact that CH_3CHF_2 did not appear to have been formed either. In this connection, it should be noted that in all of these experiments no attempt was made to control either the temperature of the reactor packing or the rate of flow of the gases through it, both of which factors would presumably have a significant influence upon the final results.

Since this reaction had yielded satisfactory results with ethane, the next logical development was obviously a study of the corresponding fluorination of methane, which had not been attempted first because preliminary experiments had indicated that it would involve complications, in spite of the simple and fundamental structure of the hydrocarbon. Accordingly, Hadley and Bigelow (40) carried out the methane fluorination, over the copper-gauze catalyst, under essentially the same conditions as have just been described, using progressively varying molar ratios of the reacting and diluent gases. It was found that when the fluorination ratio was 2:1, the dilution ratio 1:5, and thus the reaction ratio 2:1:10, the proportions of the product boiling higher than carbon tetrafluoride (−128°C.) rose to a maximum of 45 liquid volume per cent of the total condensate. It was impracticable to increase the dilution ratio still further, since under such conditions it was not easy to condense the low-boiling gases efficiently. This complex higher-boiling material, consisting of partially fluorinated methane and other products, was accumulated, passed repeatedly through alkali to remove carbon dioxide and silicon tetrafluoride, dried, and rectified in the improved Booth–Podbielniak unit. The results, to us at least, were quite surprising, since essentially no pure chemical individuals came over in the original distillations, even though some of the expected products were known to boil 30°C. apart. Instead, there were obtained in satisfactory yields from the several accumulations prepared under different operating conditions, three distinct, reproducible, heterogeneous, and therefore constant-boiling mixtures with constant molecular weights: b.p. −89°C. (mol. wt. 94), b.p. −89°C. (mol. wt. 85), and b.p. −58°C. (mol. wt. 95), together with considerable amounts of as yet unidentified products. These mixtures, either in combination or individually, were separated by freezing the lower layers, then boiling off the supernatant liquids, and finally rectifying each separately. By this means, the first mixture was shown to consist of fluoroform (b.p. −82°C.) and hexafluoroethane (b.p. −79°C.); the second of difluoromethane (b.p. −52°C.) and octafluoropropane (b.p. −36°C.); and the third of methyl fluoride (b.p. −78°C.) and hexafluoroethane (b.p. −79°C.). All of these compounds, which were already known, were satisfactorily isolated in pure condition and identified. Thus the fluorination of methane has been shown to yield, not only all of the expected fluoromethanes, but hexafluoroethane and octafluoropropane as well. This interesting result has demonstrated beyond doubt that a vapor-phase fluorination under the conditions described can not only effect the expected substitutions, but also may build up a carbon chain. This polymerization, which is associated with the free-radical or atomic nature of the reactions, has since become increasingly important.

While these experiments with hydrocarbons were going on, the fluorination of chlorinated compounds had not been entirely neglected. Fukuhara and Bigelow (35) progressively sublimed hexachlorobenzene (C_6Cl_6) from a side arm heated to 189–195°C. into a glass reaction chamber where it mingled and reacted smoothly at 55°C. and atmospheric pressure with a current of fluorine issuing through the meshes of a copper-gauze roll at the rate of about 2 liters per hour. The crude product, which was a heavy, yellow, fuming oil, condensed continuously on the walls, ran out at the bottom, and collected in a receiver below at the rate of about 8 g. per hour as long as the reactants were supplied. In general, for each 25 g. of sample consumed, 27 g. of the oil was formed; and this product, which contained roughly 13 per cent fluorine and 53 per cent chlorine, was essentially inert to fluorine, even when the gas was bubbled directly through it. The oil was considered too reactive to handle, so it was powerfully reduced by refluxing with iron powder in acetic acid solution, followed by steam distillation, in the hope of aromatizing and stabilizing it, even though a large part of the sample was destroyed in the process. The partially purified material then consisted of stable colorless liquid and some crystalline solid, which were accumulated. The oil, after washing with dilute alkali, was rectified in a high-temperature Podbielniak fractionating unit, and the solid repeatedly crystallized. As a result, no less than ten liquid and two crystalline portions were isolated, which were considered to be chemical individuals, and which possessed a number of more or less remarkable characteristics. First, all significant portions of the distillation curve, as well as the appearance of the solids, were strictly reproducible; next, all ten of the liquids, boiling constantly at various points between 32° and 75°C. at 11 mm., and the two crystalline solids (m.p. 69–70°C. and 142–143°C., respectively) contained fluorine and chlorine by analysis in atomic ratios of Cl:F corresponding to simple whole numbers; and finally, the crystals had sharp melting points, while six of the liquids had sharp freezing points, as indicated by the freezing curves, the others becoming glasses on cooling. No molecular weights, nor evidence as to the structures of the compounds, were presented.

By way of comment it may be said that it seems quite likely that the samples with both constant melting and boiling points, and Cl:F ratios as simple whole numbers as well, were actually pure chemical individuals. However, all of this material, chemically speaking, was a long way from its starting point, and probably consisted of highly halogenated, partially polymerized fragments, originally formed owing to the primary saturation and subsequent splitting of the aromatic ring by the fluorine, some of which were known to contain oxygen, and perhaps hydrogen as a result of acetolysis and hydrolysis during the stabilization process. These findings illustrate still more clearly the wisdom of studying simpler molecular species first; nevertheless the writer believes that we have shown enough to suggest that the fluorination of hexachlorobenzene offers an interesting and also a challenging field for further research.

With the preceding experience in mind, and the satisfactory technique for reacting gaseous hydrocarbons with fluorine available (see above), Calfee, Fukuhara, Young, and Bigelow (22) next studied the vapor-phase fluorination of the simple alkyl halide, ethyl chloride (b.p. +12°C.), over a copper-gauze catalyst in the manner already described. In this case, contrary to previous experience, the catalyst deteriorated rapidly unless maintained anhydrous between runs. The products were condensed by liquid air, accumulated, transferred to the Booth–Podbielniak still, and rectified in the usual way, except that the material boiling above +12°C. was withdrawn, washed with alkali, and distilled from an ordinary flask. There were formed in substantial amounts CF_4, CF_3Cl (b.p. −80°C.), CF_3CF_2Cl (m.p. −106°C., b.p. −38°C.), $CF_2{=}CCl_2$ (m.p. −116°C., b.p. 0°C.?), and CHF_2CH_2Cl (b.p. 35°C.). All of these had the correct molecular weights; the last four were analyzed and gave correct figures for both fluorine and chlorine. In addition, some evidence was found for the presence of the compounds CHF_2CHCl_2 and CH_3CHCl_2 in the complex reaction product. Locke, Brode, and Henne (55) have prepared the compound $CF_2{=}CCl_2$ by indirect methods, and after carefully checking this work, Henne and Wiest (43) have found the compound to boil at 18.9–19.0°C., a value much more nearly in line with the boiling points of related halogen compounds than 0°C. The present writer has been creditably informed that the higher value for the boiling point of $CF_2{=}CCl_2$ has recently been confirmed by others, but this work has not as yet been published. Up to the present time, the reason for the discrepancy has not been determined. In this fluorination of an alkyl chloride it is of interest that not only the expected substitution and fragmentation took place, but apparently chlorination as well; this may have been due to the presence in the reacting gases of significant amounts of chlorine fluoride, formed, presumably, by the direct displacement of chlorine by fluorine.

Since the vapor-phase fluorination of organic gases had been markedly successful, the next, and obvious, step was to adapt the process to volatile liquids in a heated reactor; so Fukuhara and Bigelow (36) fluorinated acetone, using the improved outfit illustrated to approximate scale in figure 2. This apparatus consisted essentially of the vertical, heated, screen-packed reactor (C), calibrated glass capillary (D), sump (K), heated tubes (L and L′) containing sodium fluoride to absorb hydrogen fluoride, and well-protected glass cold trap (M). In operation the sample under air pressure in A entered the reactor which was originally heated to 60°C. through D and G, while the fluorine was led in simultaneously through H and I. The more volatile products passed out at the top, were freed from hydrogen fluoride in L and L′ and finally condensed in M, while less volatile materials ran out at the bottom and were collected in K at the same time. The samples were accumulated, rectified, and finally yielded CF_4, COF_2 (b.p. −83°C.), CF_3COF (b.p. −59°C.), CF_3COCF_3 (m.p. −129°C., b.p. −28°C.), $(COF)_2$ (b.p. 26°C.), CH_2FCOCH_3 (b.p. 78°C.), and other unidentified substitution and condensation products. The known compounds COF_2, CF_3COF, and CH_2FCOCH_3 were further identified by conversion to diethyl carbonate, trifluoroacetamide, and the corresponding semicarbazone, respectively, while the new CF_3COCF_3 and $(COF)_2$ as well as the CH_2FCOCH_3 were analyzed and gave the correct figures. In addition, the hexafluoroacetone was soluble in water, stable in solution, and yielded a crystalline semicarbazone (m.p. 153°C. (decomposition)), while the oxalyl fluoride could be converted readily into the expected oxamide. It will be noticed that in this interesting reaction there were formed, not only the completely fluorinated molecule, but also the expected fragmentation products and the unexpected dimer, oxalyl fluoride. Incidentally, the boiling point of the known monofluoroacetone has been corrected. Since the procedure used was quite general, this development made possible the vapor-phase fluorination of any readily volatile organic liquid.

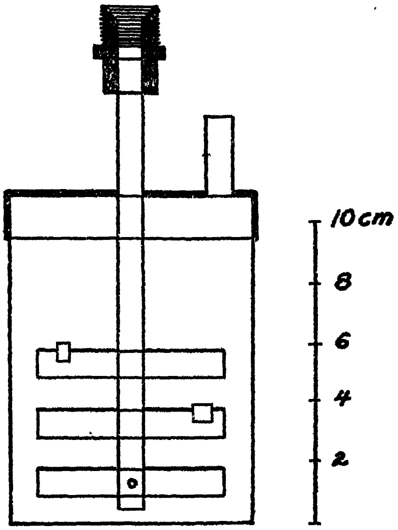


FIG. 2. Apparatus for the vapor-phase fluorination of volatile organic liquids

With a good process for the fluorination of volatile organic liquids at their disposal, Fukuhara and Bigelow (37) in 1941 next undertook a study of one of the most important of such compounds—namely, benzene itself—in spite of the fact that this hydrocarbon could scarcely be classed as having a really simple structure. They employed the same apparatus and techniques which have been described. Using a molar reaction ratio of $F:C_6H_6:N$ as 4:1:2, with the fluorine being supplied at 1.7 liters per hour and the reactor maintained at 90°C., they obtained a clean, but complex crystalline product condensed by liquid air at the rate of about 1.1 cc. per hour. This product always contained an excess of benzene, and corresponded roughly to 40 cc. of fluorinated material per 100 hr. of operation, during which time some 4–5 g. of tar was deposited inside the reactor.

A large amount of the crude product was collected and carefully rectified in Booth–Podbielniak stills, the distillation curves being strictly reproducible throughout. There resulted four pure compounds, CF_4, C_2F_6, C_3F_8, and C_4F_{10} (b.p. −2°C.), with the exception of the first, in small amounts; and four significant constant-boiling mixtures, b.p. 22°, 48°, 61–63°, and 79°C., respectively, all of which contained unreacted benzene. The first mixture was heterogeneous, carried but little hydrocarbon, and was separated by distillation with *n*-pentane. This operation brought over a new heterogeneous mixture (b.p. −5° to −3°C. at 330 mm.), the lower fluorocarbon layer of which on final rectification yielded pure saturated C_5F_{10} (b.p. 0–1°C. at 330 mm., or 22°C. at atmospheric pressure). The next mixture, also heterogeneous and obtained in the greatest amount, carried a much greater proportion of benzene, so the layers were separated mechanically. The lower, fluorocarbon layer was then distilled at 218 mm. until the contents of the column solidified, and the distillate was discarded. After this the residue, mostly solid, was further purified by distillation from trap to trap at low pressures, and finally refractionated in a special still at atmospheric pressure, yielding pure C_6F_{12} (m.p. 48–49°C., b.p. 50°C.) with a liquid range of less than 2°C. This melting point, which is somewhat difficult to attain, was extremely sensitive to the merest trace of benzene in the product. The third mixture was homogeneous, was apparently azeotropic, and contained much benzene. The sample, in combination with other similar material, was then distilled with an excess of C_6F_{12}, which brought over the heterogeneous mixture boiling at 48°C. (see above), until all of the hydrocarbon had been removed. Subsequent careful rectification of the mixture yielded pure C_6HF_{11} (b.p. 62°C.) and left a higher-boiling, unidentified residue. The last mixture was again heterogeneous, and consisted mostly of benzene. The lower, fluorocarbon layer was separated, accumulated, and rectified in a special still, yielding chiefly pure $C_{12}F_{22}$ (m.p. in bulk 19–21°C., b.p. 90°C. at 90 mm.).

Summing up, the fluorination of benzene yielded CF_4, C_2F_6, C_3F_8, C_4F_{10}, C_5F_{10}, C_6F_{12} (in greatest amount), C_6HF_{11}, and $C_{12}F_{22}$, but no aromatic fluoro compounds could be found. All of these products, except the well-known first three, were analyzed and gave the correct figures. The C_6HF_{11} and $C_{12}F_{22}$ were entirely new, and it is not at all certain that the C_4F_{10}, C_5F_{10}, and C_6F_{12} which had been described previously by Simons and Block (87) possessed either the same structures, or even the same exact compositions as ours. The results as a whole were considered to constitute strong evidence that the attack of elementary fluorine on the aromatic nucleus proceeds by an atomic chain mechanism, involving first addition, then substitution, and finally fragmentation, followed to some extent by polymerization. The detailed mechanisms discussed in this and the preceding papers will be considered more fully in Section III.

The writer cannot, in the nature of things, appraise the value of the work done under his own direction, but he would like to take this opportunity to express his gratitude for, and appreciation of the ideas, skills, perseverance, and loyalty of his student collaborators, without whom such progress as was made would not have been possible.

I. FREDENHAGEN, Cadenbach, Krefft: fluorination in the presence of a metal screen; fluorination of as-diphenylethylene

In 1934, not long after the work which has just been outlined had been started, the physical chemist Karl Fredenhagen, whose fluorine generator will be described in detail below, made a brief, but significant entrance into the field of organic fluorinations. In a paper by Fredenhagen and Cadenbach (33), in which the experimental portion is credited chiefly to the junior author, there was described for the first time a new method for moderating and controlling such reactions. The simple apparatus used consisted essentially of a horizontal, well-cooled, and rapidly rotating glass cylinder. A stationary copper inlet tube entered this rotor at one end, and extended along its central axis for about half the length of the cylinder, which was open to the air at the opposite end. The tube was closed at the inner end, but was filled with numerous perforations along its length, and wrapped with several layers of copper gauze. In operation the liquid sample was spread in a thin layer over the inner wall of the cylinder by centrifugal force, while the fluorine entered through the meshes of the gauze and reacted smoothly, without combustion, either in the vapor phase, accompanied by faint glowing, or at the large liquid surface when the sample had a low vapor pressure. In this way such compounds as benzene, toluene, paraffin oil, and palmitic acid were easily fluorinated; while no pure fluorine compounds were isolated, materials were collected which contained from 9–58 per cent of fluorine by analysis. It was of interest that when a glass inlet, even when provided with fine perforations, was used, the reactions were not moderated and consisted for the most part of combustion. Subsequently, a considerably elaborated procedure apparently based on these experiments was patented by Krefft (107), but the only definite reaction described was the conversion of $(C_6H_5)_2C{=}CH_2$ into the expected difluoride in good yield.

These experiments were clearly of the exploratory type, which so often seem to be without concrete results; yet in this case they demonstrated that when the direct fluorination of organic compounds was conducted in the presence of a metal screen, the reactions were greatly moderated in intensity and took place to a much lesser extent by combustion. This interesting observation was one of the important factors leading up to the subsequent successful development of vapor-phase fluorination in the writer's laboratory.

J. SIMONS, Block, Bond, McArthur, Herman: fluorination of carbon, carbon tetrachloride, dichlorodifluoromethane, and acetephenone

It was not until 1937 that Joseph H. Simons, a physical chemist who had been working for many years with fluorine and its compounds, began his studies dealing with direct organic fluorination. In that year Simons and Block (86) published a preliminary communication stating that the reaction of fluorine with carbon had yielded the fluorocarbons CF_4, C_2F_6, C_3F_8, C_4F_{10}, C_5F_{12}, and C_6F_{14}, all identified by their respective molecular weights only. No experimental details were given, except in regard to the fractionation, and the results were later materially modified. Subsequently, Simons and Block (87) gave a detailed report of their findings. They passed fluorine from Simons open generators (84, 85) of the Mathers type over different varieties of finely divided carbon contained in a heated copper tube. About 1 per cent of mercuric chloride or mercurous chloride was added to the charge in order to inhibit the violent explosions which otherwise took place. The exit gases were passed through an alkali solution, rigorously dried, condensed by liquid air, and finally rectified in a low-temperature fractionating unit. No definite information was given concerning the rate of flow of the fluorine or the temperature of the reactor, and it was stated that some of the cuts in the fractionation could not be made sharply. There were formed CF_4, C_2F_6, C_3F_8 (m.p. −183°C., b.p. −38°C.), two isomers of C_4F_{10} (b.p. −4.7°C. and +3.0°C., respectively), C_5F_{10} (m.p. about −12°C., b.p. 23°C.), C_6F_{12} (b.p. 51°C.), and C_7F_{14} (b.p. 80°C.), together with other unidentified fluorocarbon mixtures. The first three of these comprised 55, 13, and 8 liquid volume per cent of the crude product, respectively, while the others were present in smaller amounts, ranging from 1–5 per cent approximately. All of them had not far from the required molecular weights and, with the exception of the first two, were analyzed by decomposition over heated sodium for both carbon and fluorine. The results leave something to be desired as regards precision, but this may have been due to the analytical method, which cannot be relied upon to give precise results for carbon on account of the formation of metal carbides, as pointed out by the authors in one case. Each of the substances analyzed was new, but only the C_3F_8 can be regarded as a pure chemical individual, while the C_5F_{10} may have been approximately such. The other four, which yielded glasses on cooling, were quite likely mixtures of isomers, especially when the pyrogenetic nature of the reaction is considered. Pure C_6F_{12} has since been shown to be a white crystalline solid (m.p. 48–49°C., b.p. 50°C.) (37). Additional support for the structures claimed was given by comparing the properties of the fluorocarbons with those of related hydrocarbons, and also with those of the inert gases. Finally, a number of other physical constants pertaining to the various fractions were reported.

Later, in a paper dealing with compounds containing the CF_3 group and covering a number of both direct and indirect fluorinations, Simons, Bond, and Mc-

Arthur (88) described the fluorination of carbon tetrachloride and dichlorodifluoromethane. They passed fluorine, at the rate of about 0.16 mole per hour, through carbon tetrachloride containing 6.3 per cent by weight of arsenic in suspension and maintained at 70°C. in a special copper flask equipped with a condenser and stirrer. The reaction proceeded quietly, and the exit gases were washed with alkali and collected over water. The product contained 74 per cent of carbon tetrafluoride and only 17 per cent of chlorotrifluoromethane, as compared to 20 and 54 mole per cent of the respective gases obtained earlier by Ruff and Keim (77), using the same reagents and catalyst but somewhat different operating conditions. In the hope of increasing the proportion of chlorotrifluoromethane formed, the experiment was repeated, using only 2.3 per cent of arsenic, but the reaction was too vigorous, and was moderated by adding 2.5–3.3 per cent of bromine, which, as a more active catalyst, was expected to reduce the reaction temperature. However, the effect was greater than anticipated, and the product in this case was nearly pure trichlorofluoromethane. The desired chlorotrifluoromethane was finally obtained in high yield by passing a mixture of fluorine and dichlorodifluoromethane, with the latter in slight excess, over a mercury catalyst contained in a copper tube heated to 340–370°C. The reaction proceeded smoothly, and the product contained essentially no carbon tetrafluoride. Arsenic, silver fluoride, and bromine were ineffective as catalysts under these conditions. Neither physical constants, criteria of purity, nor yields were given in the descriptions of any of these experiments. Incidentally it was observed that the residues from the fluorination of carbon tetrachloride contained hexachlorobenzene (C_6Cl_6), which also results when carbon tetrachloride is passed through a carbon arc. This emphasized the fragmentation and subsequent polymerization which often accompany fluorinations, and also suggested that perhaps all of the hexachlorobenzene obtained in the writer's laboratory when a carbon tetrachloride solution of toluene was fluorinated (7) did not come from the hydrocarbon, as was then supposed.

Subsequently, in a paper devoted chiefly to various indirect fluorinations, Simons and Herman (89) described the action of elementary fluorine on acetophenone. They passed fluorine, diluted with five times its volume of nitrogen, at the rate of 0.2 mole per hour for 5 hr. through a solution of 0.5 mole of acetophenone in 200 g. of liquid hydrogen fluoride at 0°C. contained in a copper reaction chamber equipped with a refrigerated condenser, and so arranged that the fluorine stream was equally divided between two separate inlet tubes. The reaction, which was catalyzed by 2 g. of silver oxide, proceeded quietly, but at the end most of the fluorine passed through unreacted. Then the product was poured onto ice, extracted with ether, and the hydrogen fluoride removed by means of sodium fluoride. Finally the ether solution was dried, the solvent removed, and the product fractionated, yielding 2.1 g. (3.3 per cent) of benzoyl fluoride (b.p. 74–79°C. at 30 mm.) and 15.7 g. (20.2 per cent) of ω-difluoroacetophenone, $C_6H_5COCHF_2$. The former compound was identified by means of the corresponding benzanilide (m.p. 162–163°C.), and the latter by conversion to mandelic acid with 5 per cent sodium hydroxide solution. Neither the boiling range of the fluorinated acetophenone nor the melting point of the mandelic acid derived from it was given for this preparation, although this same halogen derivative, as prepared indirectly from the corresponding dibromo compound, was adequately characterized (b.p. 83–85°C. at 29 mm.), 2,4-dinitrophenylhydrazone (analyzed) (m.p. 221–223°C.). Also, none of the corresponding mono- or trifluoroacetophenone was formed. This unexpected reaction would appear to indicate that under mild operating conditions, a benzene nucleus may be more resistant to attack by fluorine than has hitherto been believed. Possibly the ring was deactivated by formation of the salt $(C_6H_5COHCH_3)^+F^-$. It is, however, difficult to understand how free fluorine could have passed unchanged through the system containing the acetophenone derivative unless the two were not actually in contact.

The various homologous series of fluorocarbons, the use of catalysts in direct fluorinations, and the conducting of such reactions in anhydrous hydrogen fluoride solution represent interesting and important fields of research, which merit further development.

K. MILLER: fluorination of chloroform, pentachloroethane, s-tetrachloroethane, tetrachloroethylene, and trichloroethylene (all in the liquid phase)

It was in 1940 that the writer noted with satisfaction the entrance into this field of his former student William T. Miller, now an organic chemist of standing in his own right. In a paper of merit Miller (58) described a new procedure for the direct fluorination of organic liquids. The simple apparatus, constructed of brass tubing, was roughly U-shaped, except that the vertical sides were at right angles and there was a flat middle section. The latter was to a large extent divided into upper and lower compartments by a horizontal baffle plate, at one end of which was a small vertical well containing a mechanical stirrer. This was activated through one of the arms of the U and so arranged as to produce a definite flow of liquid around the ends of the baffle and in opposite directions in the two compartments. The arm which carried the vertical stirrer shaft was tightly closed at the top, except for a small exit tube, while the other one carried the fluorine line, which entered through a rubber stopper protected by thin copper foil and was tightly wrapped with a roll of copper gauze at the inner end. In operation, the lower section of the apparatus was filled with the liquid organic sample, while the undiluted fluorine, substantially free from oxygen (60), entered through the wire gauze at the rate of 1.7 liters per hour (5 amp.), bubbled progressively in counter-current flow through the sample, and finally, if unreacted, passed out together with any volatile products through the exit tube adjacent to the stirrer shaft. This method resulted in excellent mixing and smooth reactions, taking place either in the vapor phase or at the liquid surface. The organic samples were always in large excess, and finally served as solvents for the products. These mixtures were in all cases carefully washed, dried, and then repeatedly rectified.

Under these conditions, the fluorination of chloroform ($CHCl_3$) for 15 amp.-hr. at 0°C. gave a crude product essentially 100 per cent by weight of the original sample, which on fractionation yielded trichlorofluoromethane (b.p. 23–25°C.) equivalent to 5.6 per cent of the crude, and hexachloroethane (m.p. 184°C. (subl.)), 0.5 per cent, as residue after removal of the excess of chloroform. It was observed that this material was too volatile for smooth fluorination by this method, and could be handled more efficiently in a packed tube.

In a like manner pentachloroethane (CCl_3CHCl_2), fluorinated for 50 amp.-hr. at 90°C. ± 3°, reacted smoothly and gave 85 per cent of its weight of crude product, which on rectification yielded $CFCl_2CFCl_2$ (m.p. 22–24°C., b.p. 90–92°C.), 1.7 per cent of the crude; $CCl_2{=}CCl_2$ (b.p. 117.5–118°C.), 5.2 per cent of the crude; CCl_3CFCl_2 (m.p. (redist.) 99.8–100°C., b.p. 137–139°C.), 28 per cent of the crude; and C_2Cl_6 (m.p. 183–184°C. (recryst., subl.)), 9.5 per cent of the crude, together with much unreacted sample. By extraction of the still residue with alcohol, and subsequent crystallization by slow evaporation from glacial acetic acid and formic acid (1:1), there was also obtained analytically pure $CCl_3CCl_2CCl_2CCl_3$ (m.p. 80–81°C.). Similarly, *s*-tetrachloroethane ($CHCl_2$-$CHCl_2$) when reacted for 70 amp.-hr. at 50°C. ± 2° gave 95 per cent of its weight of crude product, which on redistillation yielded $CCl_2{=}CHCl$ plus $CFCl_2$-$CFCl_2$ (b.p. 90–95°C. ?); $CHCl_2CFCl_2$ (b.p. 115.5–116.5°C.), 22 per cent of the crude (redistilled, b.p. 115.7°C.); and CCl_3CHCl_2 (b.p. 159–160°C.), 3.6 per cent of the crude; together with a large amount of unchanged material. The $CCl_2{=}CHCl$ and $CFCl_2CFCl_2$ could not be separated by distillation, but the former was identified as $Hg(CCl{=}CCl_2)_2$ (m.p. 82–83°C.) and the latter was finally obtained slightly impure (m.p. 19–23°C., b.p. 92–93°C.). No hexachloroethane could be isolated from the reaction mixture. Likewise, tetrachloroethylene ($CCl_2{=}CCl_2$) after reaction for 60 amp.-hr. at 0°C. gave 106 per cent of its weight of crude product, which on fractionation yielded $CFCl_2CFCl_2$ (b.p. 89–94°C.) (redistilled, m.p. 24.43–24.73°C. (corr.), b.p. 92.3°C.), 16 per cent of the crude; CCl_3CFCl_2 (b.p. 136–138°C.) (redistilled, m.p. 99–99.5°C. (in a sealed tube)), 14 per cent of the crude; and $CFCl_2CCl_2CCl_2CFCl_2$ (m.p. 4–5°C., b.p. 152.5°C. at 20 mm.), 19 per cent of the crude; along with the unchanged sample.

Next, trichloroethylene ($CCl_2{=}CHCl$), when fluorinated for 50 amp.-hr. at 0°C., gave 102 per cent of its weight of crude product, which on rectification yielded at first a very volatile portion, b.p. up to 40°C., which apparently contained $CFCl{=}CHCl$, identified as $CFClBrCHClBr$ (b.p. 160–162°C.). Then came a mixture (b.p. 71.5–73°C.) presumably consisting of $CFCl_2CHFCl$ (known b.p. 72.5°C.) and $CFCl{=}CCl_2$ (known b.p. 72.1°C.), 4.3 per cent of the crude, which could not be separated, but which yielded on chlorination both $CFCl_2CFCl_2$ and CCl_3CFCl_2. This was followed by another mixture (b.p. 114–116°C.), probably $CHFClCCl_3$ (known b.p. 117°C.) and $CHCl_2CFCl_2$ (known b.p. 116.6°C.), 5.4 per cent of the crude, which could not be separated but which, on chlorination, yielded only CCl_3CFCl_2. Further fractionations yielded a portion (b.p. 122.0–125.5°C. at 25 mm.), 12 per cent of the crude, which on repeated crystallization finally gave analytically pure $C_4H_2Cl_6$ (m.p. 9.5–11°C., b.p. 125.5°C. at 25 mm.), which would not add bromine and was not $CHCl_2CCl{=}CClCHCl_2$. From the still residues there was obtained some solid, which on repeated crystallizations finally yielded analytically pure $C_4H_2Cl_8$ (m.p. 75–76°C.), which was probably not $CHCl_2CCl_2CCl_2CHCl_2$. Finally, $CCl_2{=}CCl_2$ and $CCl_2{=}CHCl$ were fluorinated in CF_3CCl_3 (Freon 113) solution at 0°C. In the first case, the results were similar to those given above, and in the second, only one analytically pure product was isolated, which was $C_4H_2F_2Cl_6$ (m.p. 55–56°C.).

The work which has just been described very briefly represents the first successful attempt to react elementary fluorine with pure liquid organic compounds containing more than one carbon atom. The contribution is definite, and the reactions, which take place under conditions favoring an atomic mechanism, will be discussed further in Section III. A study of this paper reveals throughout a degree of accuracy and precision which creates confidence in the results.

L. Patents

In addition to the work which has been detailed or referred to in the foregoing pages, a number of patents have been taken out in recent years, covering various processes involving direct fluorination. A number of these which have been abstracted will be summarized briefly below, but it should be understood that no exhaustive patent search has been attempted.

In 1934 the du Pont Company (98) patented abroad a process for fluorinating organic compounds in inert solvents, such as hydrogen fluoride, dichlorodifluoromethane, trichlorotrifluoroethane ($C_2F_3Cl_3$), and others, in the presence of catalysts such as iodine or antimony pentachloride. Benzotrifluoride, under these conditions, underwent polymerization. A year later I. G. Farbenindustrie (101) listed a procedure for carrying out such fluorinations in a chamber divided into small compartments by an inert packing such as calcium fluoride. Following this, Calcott and Benning (to du Pont) (94) and also Daudt and Parmelee (to du Pont) (96) took out American patents for direct fluorination in solvents, including many objects, conditions, and claims, but the procedures as described did not lead to the isolation of pure chemical individuals. Henne (to General Motors) (105) patented a process for reacting aliphatic halides containing halogens other than fluorine with elementary fluorine in the presence of pentavalent antimony compounds as catalysts. At least one halogen atom was replaced by fluorine, and the products separated in a dephlegmator. Calcott (to du Pont) (93) has also described a process for making a colored pigment by the direct fluorination of a copper phthalocyanine in liquid hydrogen fluoride. Finally, patents have been granted to I. G. Farbenindustrie (102, 103, 104) for a process involving the direct fluorination of oil- or wax-impregnated fibrous materials to be used for packaging or rubber for cell insulators; and to Gaylor (to Standard Oil Development) (99, 100) for a method of making fluorinated oils suitable for coating or impregnating wood and cloth, or as lubricating oils, antioxidants, and fireproofing agents. In the opinion of the writer, the procedures described in these patents are all characteristically indefinite, with the result that it is not possible to determine from the published reports to what extent they may lead to valuable developments.

This concludes the factual summary of the work in this field which has been published to the end of the year 1945.

M. Tabulation of fluorine compounds

Table 1 presents a list of fluorine compounds as prepared for the first time by direct fluorination.

III. THEORETICAL CONSIDERATIONS

The several investigators who have studied the action of elementary fluorine on organic compounds and whose work has been detailed in the foregoing pages have spent the greater part of their time and energies in developing experimental methods and accumulating data rather than in postulating and discussing reaction mechanisms. The reasons for this are not far to seek, since the field is new, the operating techniques difficult, and, until very recently, the literature background of necessity very scant. However, a number of theoretical discussions have been presented by some of the more recent investigators and these will be outlined in the same order as that used in the preceding section.

A. Contributions of Ruff

Ruff and Bretschneider (71, 74), it will be recalled, have shown that graphite or Norite will react with fluorine below the ignition point to form a very interesting compound best represented by the formula $(CF)_x$. With the added collaboration of Ebert, they have postulated a detailed mechanism for the formation of this substance, which was considered to be in substantial accord with the physical and chemical properties of the compound. The essentials of this theory are as follows: The fluorine molecules, first attracted to the carbon surface by van der Waals forces, next penetrate into the graphite lattice between the base planes of carbon rings of which it is composed. As a result, the distance between the planes is expanded from 3.4 A. to 8.17 A. and the fluorine molecules are changed to ions, presumably by acquiring electrons from the carbon rings. Then these ions, with a radius of 1.33 A., are packed in the most efficient possible manner into the volume provided by the expanded lattice, which, according to these investigators, requires them to lie in an orderly pattern in no less than six parallel planes, intermediate between the carbon base planes. The arrangement also necessitates that alternate carbon base planes be slightly distorted, by the displacement of certain carbon atoms in opposite directions. It was found that the volume necessary to pack in one fluoride ion per carbon atom in this manner required a lattice expansion which in turn corresponded very closely to the observed drop in density during the fluorination process. Also, the loss of electrons by the carbon rings to the fluorine should break up the conjugated systems in these rings, and so destroy the electrical conductivity of the graphite, which was found to be the case. Further, this arrangement assumes that while the distance between the carbon base planes is greatly increased, the ring structure within these planes is not greatly altered, and these assumptions were well supported by the observed x-ray data. Finally, if such a structure were rapidly

TABLE 1

Fluorine compounds as prepared for the first time by direct fluorination, either in analytically pure condition or otherwise adequately characterized

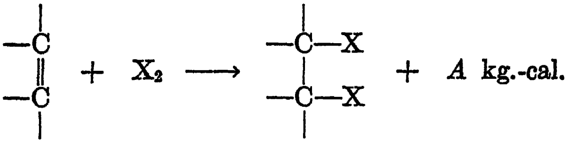
COMPOUND	MELTING POINT	BOILING POINT	INVESTIGATOR	REFERENCE	REMARKS
	°C.	°C.			
CH_3F		−78	Bigelow	(40)	
CH_2F_2		−52	Bigelow	(40)	
CHF_3		−82	Bigelow	(40)	
CF_4	−186	−130	Lebeau and Damiens	(53, 54)	
			Ruff	(76)	
$CFCl_3$	−111	24.1	Ruff	(77)	
CF_2Cl_2	−160	−28	Ruff	(77)	
CF_3Cl	−181	−80	Ruff	(77)	
CF_3NO	−150?	−84	Ruff	(80)	Mixed with $FCONF_2$
$FCONF_2$	−152.5	−82.4	Ruff	(80)	
CHF_2CH_2F	−84	3	Bigelow	(92)	
CHF_2CHF_2		−23.5 to −22.5	Bigelow	(92)	
CF_3CHF_2	−103	−48.5	Bigelow	(92)	
C_2F_6	−101	−78	Bigelow	(20, 73)	Originally prepared indirectly by Ruff (73)
CHF_2CH_2Cl		35	Bigelow	(22)	
$CHCl_2CFCl_2$		115.7	Miller	(58)	
CCl_3CFCl_2	100		Bockemüller	(10)	
$CFCl_2CFCl_2$	26.5	91	Bockemüller	(10, 44)	Highest purity: m.p. 28.1°C., b.p. 91.5°C.
CF_3CF_2Cl	−106	−38	Bigelow	(22)	
$(COF)_2$		26	Bigelow	(36)	
$(CNF_3)_2$	−128	−40	Ruff	(79)	Mixture of at least five isomers
C_3F_8	−183	−38	Simons	(87)	
CH_2FCOCH_3		78	Bigelow	(36)	
CF_3COCF_3	−129	−28	Bigelow	(36)	
C_4F_{10}			Simons	(87)	Reported properties conflicting (see text above)
			Bigelow	(37)	

heated, the increase in volume produced by the formation of carbon tetrafluoride within the expanded lattice would be expected to result in the explosive disruption of the material, a result which actually took place under such conditions. From all this it will be clear that the theoretical assumptions involved have considerable experimental support.

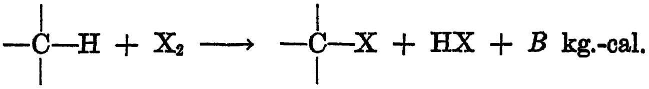
The writer has been privileged to discuss the speculations just outlined with Drs. P. M. Gross, F. London, and M. E. Hobbs of this Department, who have given material assistance which is hereby gratefully acknowledged. While there is no doubt that the experimental work on which the theory is based is thoroughly sound, none of us can agree with the proposed ionic nature of the lattice of $(CF)_x$. Fluoride ions are negatively charged, and if they were packed together in successive layers as proposed, the repulsive forces would exceed the cohesive forces, and the whole structure would fly apart and so could not exist. Dr. London has proposed another possible structure involving only covalent links, which would greatly reduce the repulsive forces. In this model, the fluorine atoms are arranged in an orderly pattern in two parallel planes, one above and the other below each carbon base plane. The fluorine atoms are so arranged that two of them, one above and the other below, are connected by covalences to each carbon ring, making use of the two available electrons possessed by each such ring due to its partially unsaturated character. If the normal carbon–fluorine distance of approximately 1.4 A. is assumed, this structure requires that the adjacent fluorine planes be separated by a distance of about 5 A., in order to account for the observed density value. At first glance this distance would seem to be excessive, but such an equilibrium may readily be explained in the following manner. The electronegative fluorine atoms, although not ionized, are in any case partially charged and the adjacent layers of these atoms would be subject to fairly strong electrostatic repulsive forces, which could be adequately counterbalanced by the available long-range van der Waals attractive forces only at relatively large distances. These partially charged fluorine atoms represent an intermediate state between ions whose charges would exert repulsive forces too great to be counterbalanced at such distances as would be compatible with the density of the compound, and neutral fluorine molecules, as in liquid fluorine, which are subject to little or no electrostatic repulsion, but are kept apart by short-range van der Waals repulsive forces only. Altogether, the proposed concept represents a stable structure, accounts well for the physical and chemical properties of the compound, and last, but not least, seems much more satisfying than the one postulated by Ruff and his collaborators.

B. Contributions of Bockemüller

Bockemüller (10) in his theoretical discussion has called attention to the fundamental energy relationships involved in organic fluorinations, and has shown how they may readily account for the powerful and even "catastrophic" reactions which have often been observed. For example, in the typical addition and substitution reactions



$$\begin{array}{c}|\\ -\mathrm{C}\\ \|\\ -\mathrm{C}\\ |\end{array} + X_2 \longrightarrow \begin{array}{c}|\\ -\mathrm{C}-\mathrm{X}\\ |\\ -\mathrm{C}-\mathrm{X}\\ |\end{array} + A\ \text{kg.-cal.} \qquad (1)$$



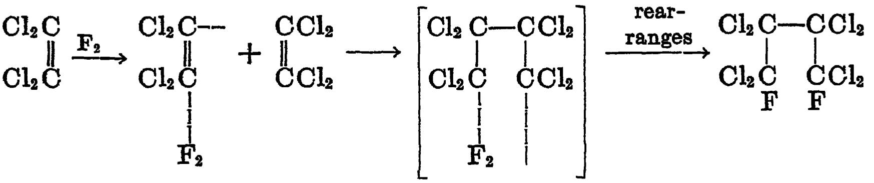
$$-\overset{|}{\underset{|}{\mathrm{C}}}-\mathrm{H} + X_2 \longrightarrow -\overset{|}{\underset{|}{\mathrm{C}}}-\mathrm{X} + \mathrm{HX} + B\ \text{kg.-cal.} \qquad (2)$$

his values for A and B are approximately 107 and 103 when $X = F$, while they are only 33 and 23, respectively, when $X = Cl$. These heats of reaction are to be compared with the heat of formation (or disruption) of a C—C bond, which is about 71 kg.-cal. From the comparison it is obvious that the over-all heats of addition or substitution in the case of fluorine, unless adequately dissipated, are ample to disrupt an adjacent C—C linkage; furthermore, if the considerable heats of activation also set free at the moment of molecular or atomic impact are taken into consideration as well, the observed combustions, charring, and polymerizations may readily be explained.

It was on the basis of these considerations that Bockemüller decided to carry out fluorinations in dilute carbon tetrachloride solutions, in which the reacting molecules would be surrounded by layers of presumably inert solvent, and to use fluorine diluted with carbon dioxide, where the halogen itself was surrounded by inert gas molecules as well. These conditions, together with the relatively low temperatures maintained (0–15°C.), favored the rapid dissipation of the energy released by the reactions. On the other hand, since fluorine is but slightly soluble in carbon tetrachloride, the reactions must have taken place at the vapor–liquid interface, and furthermore, fluorine has been shown to react with carbon tetrachloride even at 0°C., at least in the presence of a chlorine acceptor (7), both of which factors would be disadvantageous. Nevertheless quite a number of saturated and unsaturated aliphatic compounds were fluorinated successfully by this method, as detailed in the preceding section, and this success was presumably due to relatively favorable rates of reaction prevailing under the specified operating conditions.

Among other things this investigator reported that the fluorination of butyric or isobutyric acid under the conditions just mentioned yielded either β- or γ-substitution products, but none of the corresponding α-derivatives. From this evidence, which was well supported, he drew the obvious conclusion that the mechanism of α-substitution must be different from that of β- and γ-substitution, but did not discuss the matter further. This interesting observation, once thought to be more or less incredible, fits in well with modern ideas, since Kharasch and Brown (47) have recently shown that when *n*-butyric acid or the corresponding acid chloride interacted with an *atomic* reagent, such as sulfuryl chloride, in the presence of benzoyl peroxide, β- and γ-substitution took place almost exclusively, whereas the usual catalyzed α-substitution is generally regarded as an *ionic* change. Thus it appears, perhaps for the first time, that direct fluorinations in the aliphatic series, even in dilute solutions, proceed to a great extent at least by atomic chain mechanisms, a fact which has since been well substantiated, as will appear below.

It was also found that the direct fluorination of tetrachloroethylene (C_2Cl_4) yielded not only the expected $CFCl_2CFCl_2$, but also considerable $CFCl_2CCl_2$-CCl_2CFCl_2 and some CCl_3CFCl_2. The formation of the dimer was explained by assuming the interaction of an original, unstable addition product with a second molecule of C_2Cl_4 as follows:

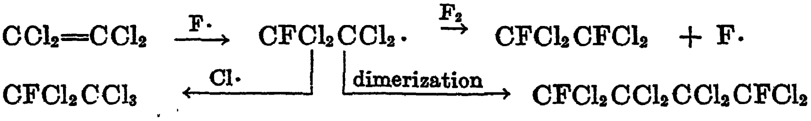


(3)

and the presence of the pentachloro compound was accounted for by the simultaneous addition of fluorine and chlorine to the original substance. It was assumed, naturally, that the necessary chlorine had been set free by a simultaneous side reaction.

The entire change may readily be explained on the basis of the more modern atomic mechanism according to the scheme

$$F_2 \rightleftarrows 2F\cdot \quad \text{(thermal or catalytic)} \tag{4}$$



$$\begin{tikzcd}
CCl_2{=}CCl_2 \arrow[r, "F\cdot"] & CFCl_2CCl_2\cdot \arrow[r, "F_2"] \arrow[dl, "Cl\cdot"'] \arrow[dr, "\text{dimerization}"] & CFCl_2CFCl_2 + F\cdot \\
CFCl_2CCl_3 & & CFCl_2CCl_2CCl_2CFCl_2
\end{tikzcd}$$

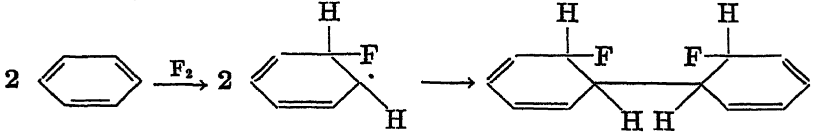
(5)

the necessary chlorine atom having been set free by the reaction

$$RCl + F\cdot \rightarrow RF + Cl\cdot \tag{6}$$

This assumes, of course, that the chlorine atom collides with the excess of organic radical before either reacting with fluorine to form chlorine fluoride, or dimerizing to give molecular chlorine. Three other possibilities—namely, that either molecular fluorine or chlorine fluoride might have reacted directly with the haloölefin, or that the organic radical could have attacked chlorine molecules—are not necessarily excluded.

Finally, Bockemüller studied the fluorination of a number of aromatic compounds, and found that simple substitution did not take place, but that addition to the nucleus occurred instead and ultimately resulted in the formation of highly polymerized products containing fluorine, according to the scheme:



(7)

which could be repeated indefinitely. This formulation is essentially the same as the corresponding atomic mechanism, in which a fluorine atom would be represented as attacking a nuclear double bond, followed by dimerization of the resulting radical. Bockemüller then considered that if the unsaturated intermediates could be quickly saturated during the fluorination, the undesirable polymerization might be prevented, and definite products formed. Accordingly, he carried out the fluorination of benzoic acid in dilute carbon tetrachloride solution in the presence of a considerable concentration of bromine, which itself does not react appreciably with either the fluorine or the acid under these conditions. The resulting product contained practically no polymeric material, and finally yielded for the most part two portions having closely the compositions required by the formulas $C_6H_6F_4Br_2$ and $C_6H_6F_3Br_3$, which were thought to be derivatives of cyclohexane. This evidence, although not entirely conclusive, represents perhaps the first attempt to demonstrate experimentally the now well-recognized fact that the primary attack of elementary fluorine on the aromatic nucleus is one of addition by an atomic mechanism.

C. Contributions of Bigelow

The writer and his students have also been convinced for a long time that direct fluorinations take place, either in the vapor phase or in solution, by atomic chain mechanisms, and they have advanced a number of such postulations, as well as the more classical reactions, in order to explain the observed facts. Admittedly, these speculations lack rigid proof in all cases, but they have been put forward first, because they correlate and clarify most of the facts logically from a modern viewpoint, and further, since speculation, understood to be such, even if later disproved, is thought provoking and may lead to progress.

Young, Fukuhara, and Bigelow (92) fluorinated ethane in the vapor phase over a copper-gauze catalyst and obtained CF_4, C_2F_6, CH_3CHF_2, CHF_2CHF_2 and CHF_2CH_2F, but no mono- or difluoroethane. At this time they proposed the classical reactions to explain this, as follows:

$$C_2H_6 \xrightarrow{F_2} C_2H_5F \xrightarrow{-HF} CH_2{=}CH_2 \xrightarrow{F_2} CH_2FCH_2F \xrightarrow{-HF} CH_2{=}CHF \xrightarrow{F_2} CH_2FCHF_2 \quad (8)$$

followed by straight substitution and finally cleavage. This interpretation was based on the known instability of CH_3CH_2F and CH_2FCH_2F at elevated temperatures. This scheme, however, is not in line with the newer view that if the life of an intermediate is so short that the compound never exists in sufficient concentration to be isolated, it is better to assume that it is not formed at all, provided an adequate mechanism is available. If we postulate the formation of fluorine atoms first, according to equation 4 above, then we may write

$$CH_3CH_3 + F\cdot \rightarrow HF + CH_3CH_2\cdot \quad (9)$$

This radical should easily lose a hydrogen atom to fluorine

$$CH_3CH_2\cdot + F_2 \rightarrow HF + F\cdot + CH_2{=}CH_2 \quad (10)$$

followed by

$$CH_2{=}CH_2 + F\cdot \rightarrow \cdot CH_2CH_2F \tag{11}$$

or perhaps reactions 10 and 11, being practically simultaneous, could be combined to give

$$CH_3CH_2\cdot + F_2 \rightarrow HF + \cdot CH_2CH_2F \tag{12}$$

thus eliminating ethyl fluoride and ethylene from the picture. Then

$$\cdot CH_2CH_2F + F_2 \rightarrow HF + CH_2FCHF\cdot \tag{13}$$

similarly avoiding ethylene fluoride and vinyl fluoride. This last radical presumably would not be expected to lose a hydrogen atom to fluorine as readily as its predecessors, but rather to react in the more conventional manner:

$$CH_2FCHF\cdot + F_2 \rightarrow F\cdot + CH_2FCHF_2 \tag{14}$$

CH_2FCHF_2 was the least fluorinated molecule actually isolated. It should be noted here that in the series 9, 12, 13, and 14, only equation 9 requires active fluorine atoms, which are present in far lower concentrations than the corresponding molecules. The compound CH_2FCHF_2 contains the stable, resonating $>CF_2$ group (15) and, once formed, would undoubtedly require an active atom for further fluorination; as these are in scarce supply, it accumulates in sufficient amount to be isolated. Then we have

$$CH_2FCHF_2 + F\cdot \rightarrow HF + \cdot CHFCHF_2 \tag{15}$$

$$\cdot CHFCHF_2 + F_2 \rightarrow F\cdot + CHF_2CHF_2 \tag{16}$$

and so on until hexafluoroethane is finally reached, and ultimately cleaved to carbon tetrafluoride. Naturally any such scheme, although it accounts for the observed results, cannot represent more than a part of the complex changes occurring in such a fluorination mixture. It also implies the formation of dimers and polymers from these radicals, which were not isolated; but their formation is by no means excluded, since the complex reaction product contained much material in transition fractions, residues, and the like, which could not be separated and identified.

Hadley and Bigelow (40) later fluorinated methane in a similar manner, and obtained not only the expected methyl fluoride, difluoromethane, fluoroform, and carbon tetrafluoride but also hexafluoroethane and octafluoropropane, thus demonstrating conclusively that a carbon chain may be built up as well as broken down by the action of elementary fluorine. All of these results may be explained very well by means of atomic mechanisms even simpler than those just outlined. For example, we have

$$CH_4 + F\cdot \rightarrow HF + CH_3\cdot \tag{17}$$

$$CH_3\cdot + F_2 \rightarrow F\cdot + CH_3F \tag{18}$$

and so on until carbon tetrafluoride is finally produced. However, when

$$CHF_3 + F\cdot \rightarrow HF + CF_3\cdot \tag{19}$$

the highly fluorinated radical does not react quite so readily with fluorine molecules, and so part of it dimerizes instead:

$$2CF_3\cdot \rightarrow C_2F_6 \qquad (20)$$

Finally the following changes also appear to be reasonable:

$$CH_2F_2 + F\cdot \rightarrow HF + CHF_2\cdot \qquad (21)$$

$$CF_3\cdot + CHF_2\cdot \rightarrow CF_3CHF_2 \qquad (22)$$

$$CF_3CHF_2 + F\cdot \rightarrow HF + CF_3CF_2\cdot \qquad (23)$$

$$CF_3\cdot + CF_3CF_2\cdot \rightarrow C_3F_8 \qquad (24)$$

Although no higher polymers were isolated, these might well have been present in the highly complex product.

Continuing these studies, Fukuhara and Bigelow (36) turned to the fluorination of the volatile liquid acetone and finally obtained CH_2FCOCH_3, CF_3COCF_3, CF_3COF, $(COF)_2$, COF_2, and CF_4, together with other unidentified polyfluoroacetones and other products. Here it will be observed that fragmentation products instead of polymerization products have become important. However, the atomic mechanism seems quite adequate to account for all the known results in this instance also, as follows:

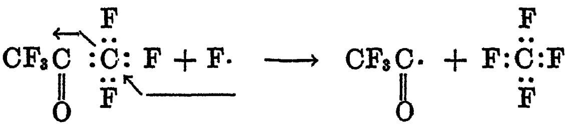
$$CH_3COCH_3 + F\cdot \rightarrow HF + CH_3COCH_2\cdot \qquad (25)$$

$$CH_3COCH_2\cdot + F_2 \rightarrow F\cdot + CH_3COCH_2F \qquad (26)$$

and so on until the completely fluorinated compound CF_3COCF_3 is reached. After that, we have

$$CF_3COCF_3 + F\cdot \rightarrow CF_3CO\cdot + CF_4 \qquad (27)$$

or, more explicitly,



$$\underset{\overset{\|}{O}}{CF_3C}\ \overset{F}{\underset{\ddot{F}}{:\ddot{C}:}}F + F\cdot \longrightarrow \underset{\overset{\|}{O}}{CF_3C}\cdot + F:\overset{F}{\underset{F}{\ddot{\underset{\cdot\cdot}{C}}}}:F \qquad (28)$$

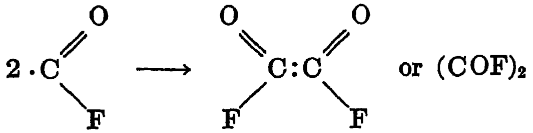
meaning that the unpaired electron on the fluorine atom enters the octet of a terminal carbon atom on the ketone, in order to form a pair. The resulting complex (see, for example, reference 16), of short life, then rapidly dissociates yielding the stable CF_4 and the $CF_3CO\cdot$ radical, thus completing the fragmentation of the carbon chain. Theoretically the fluorine atom could have attacked the ketone at several other points, but such reactions not only appear improbable, but also would lead to the formation of products which were not actually observed. The formation of the $CF_3CO\cdot$ radical moreover, accounts perfectly for all the other compounds isolated, as follows:

$$CF_3CO\cdot + F_2 \rightarrow F\cdot + CF_3COF \qquad (29)$$

$$CF_3COF + F\cdot \rightarrow \cdot COF + CF_4 \qquad (30)$$

Now the radical ·COF may either react with fluorine or dimerize:

$$\cdot COF + F_2 \longrightarrow F\cdot + COF_2 \qquad (31)$$

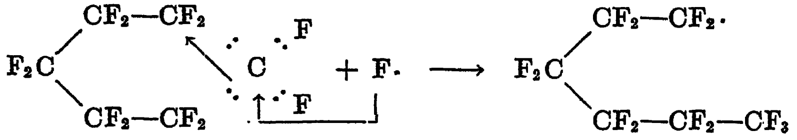


(32)

It is of interest that Kharasch and coworkers (48) have found positive evidence for the existence of the analogous ·COCl radical. Finally, this mechanism would predict the formation of the dimer $CF_3COCOCF_3$, and while this compound was not isolated, definite evidence indicating its presence in the reaction product was reported.

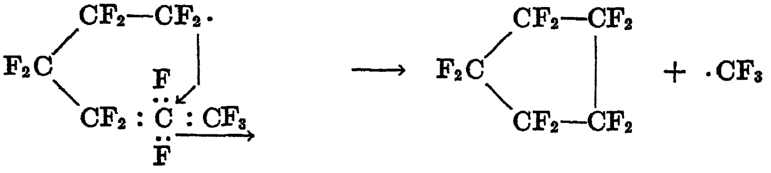
The next logical step in this series seemed to be the direct fluorination of benzene, and so Fukuhara and Bigelow (37) carried out this reaction, which proceeded easily and led to a clean but very complex product. Ultimately there were obtained C_6F_{12}, C_6HF_{11}, $C_{12}F_{22}$, C_5F_{10}, C_4F_{10}, C_3F_8, C_2F_6, and CF_4, but no aromatic substitution products whatever. These results indicate strongly that the reaction proceeded by an atomic mechanism involving, first addition, then substitution, and finally fragmentation, which may have been followed later by dimerization or polymerization of the fragments. There is no definite evidence to show that these processes do not go on simultaneously, but it seems more probable that they take place for the most part successively. No intermediates between benzene itself and C_6HF_{11} could be isolated under the operating conditions used, and need for further work in this important field is clearly indicated. However, any possible saturated intermediate of the $C_6H_6F_6$ type must have undergone direct substitution rather than thermal loss of hydrogen fluoride, since in the latter case aromatic fluorobenzenes should have been formed, but were not found.

In accord with the atomic mechanism already postulated, the formation of C_5F_{10} by fragmentation may be explained in the following manner:



(33)

followed by



(34)

and

$$2CF_3\cdot \longrightarrow C_2F_6 \qquad (35)$$

This scheme represents the opening of the six-membered ring by an active fluorine atom (33), followed by the closing of the perhaps even more stable five-membered ring (34), through an attack by the unpaired electron on the octet of the spatially adjacent $>CF_2$ group with the simultaneous expulsion of a $\cdot CF_3$ radical from the molecule. Since the fluorine was always at relatively low concentration during the reaction, this hypothesis appears quite plausible and accords well with the observed fact that neither of the acyclic fluorocarbons C_6F_{14} or C_5F_{12} could be isolated from the products. It should be recognized here, however, that all such theories depend for their validity on the assumptions that the molecules in question were actually cyclic, and that they are really derivatives of cyclohexane and cyclopentane, respectively. These assumptions are certainly likely, especially since C_6F_{12} was the major product of the fluorination and neither fluorocarbon showed any signs of unsaturation, but they still lack rigid proof. One further indication that the C_6F_{12} had a symmetrical structure was its relatively high melting point, 48–49°C. Also, it is possible, though less likely, that C_5F_{10} could have been produced by polymerization of smaller fragments rather than by degradation of the larger molecules. All of the fluorocarbons of lower molecular weight which were obtained may readily be accounted for similarly, either by fragmentation or polymerization, or by both.

Finally, the formation of the interesting diphenyl derivative $C_{12}F_{22}$ may be explained without difficulty as a dimerization in the following manner:

$$C_6HF_{11} + F\cdot \rightarrow HF + C_6F_{11}\cdot \qquad (36)$$

$$2C_6F_{11}\cdot \rightarrow C_{12}F_{22} \qquad (37)$$

In the vapor phase, and in the absence of an excess of fluorine molecules or atoms' this relatively stable radical, which could not disproportionate easily, might well be expected to dimerize. Even if it were to attack a hydrocarbon molecule, the result would be the formation of C_6HF_{11}, and the process would then be repeated. Altogether, it appears that essentially all the known facts concerning this interesting and important fluorination may be correlated and clearly explained on the basis of an atomic mechanism.

D. Contributions of Miller

In concluding this brief theoretical survey, mention should be made of the recent interesting work by Miller (58), dealing with the direct fluorination of liquid organic halogen compounds. He considered that these reactions, which took place either at a vapor–liquid interface or in the gas phase (see above), proceeded essentially by atomic mechanisms, but did not outline any of them specifically in his published paper, which was devoted in the main to the presentation of precise experimental results. The present writer has therefore selected two of these liquid fluorinations as illustrations to show how well the atomic mechanism may be applied to correlate and explain the results obtained under these conditions.

Miller found that the fluorination of CCl_3CHCl_2 yielded, ultimately, in the order of their boiling points, $CFCl_2CFCl_2$, $CCl_2{=}CCl_2$, CCl_3CFCl_2, CCl_3CCl_3, and $CCl_3CCl_2CCl_2CCl_3$. The formation of these compounds may be explained as follows:

$$CCl_3CHCl_2 + F\cdot \rightarrow HF + CCl_3CCl_2\cdot \quad (38)$$

$$CCl_3CCl_2\cdot + F_2 \rightarrow F\cdot + CCl_3CFCl_2 \quad (39)$$

$$CCl_3CCl_2\cdot + F_2 \rightarrow F\cdot + ClF + CCl_2{=}CCl_2 \quad (40)$$

$$2CCl_3CCl_2\cdot \rightarrow CCl_3CCl_2CCl_2CCl_3 \quad (41)$$

$$CCl_3CHCl_2 + ClF \rightarrow HF + CCl_3CCl_3 \quad (42)$$

$$CCl_2{=}CCl_2 + F_2 \rightarrow CFCl_2CFCl_2 \quad (43)$$

It is clear, of course, that such a complex mixture could be formed by many different routes, of which the above seem most logical since they require a minimum of the activated fluorine atoms, which are present at low concentrations. The compound chlorine fluoride is used as the chlorinating agent on account of the very high heat of formation of hydrogen fluoride, and a considerable concentration of this substance can easily result from the impact of fluorine atoms on any chlorinated molecule according to the scheme:

$$RCl + F\cdot \rightarrow ClF + R\cdot \quad (44)$$

It is not considered probable that significant concentrations of either molecular or atomic chlorine were present in the reaction mixture, although the latter could have been formed according to the equation

$$RCl + F\cdot \rightarrow RF + Cl\cdot \quad (45)$$

or even

$$R_2C{=}CR_2 + FCl \rightarrow R_2CFCR_2\cdot + Cl\cdot \quad (46)$$

and so make possible a certain amount of direct chlorination.

Miller also reported that the fluorination of $CHCl_2CHCl_2$ yielded $CFCl_2CFCl_2$, $CCl_2{=}CHCl$, $CHCl_2CFCl_2$, and CCl_3CHCl_2, but no C_2Cl_6. These results may be accounted for by the changes:

$$CHCl_2CHCl_2 + F\cdot \rightarrow HF + CHCl_2CCl_2\cdot \quad (47)$$

$$CHCl_2CCl_2\cdot + F_2 \rightarrow F\cdot + CHCl_2CFCl_2 \quad (48)$$

$$CHCl_2CCl_2\cdot + F_2 \rightarrow F\cdot + ClF + CHCl{=}CCl_2 \quad (49)$$

This implies that the residual hydrogen atom is well shielded by the large adjacent chlorine atoms. It is also quite possible that this last change is entirely thermal, and requires no fluorine, as

$$CHCl_2CCl_2\cdot \xrightarrow{\text{heat}} CHCl{=}CCl_2 + Cl\cdot \quad (50)$$

Then

$$CHCl_2CHCl_2 + ClF \rightarrow HF + CCl_3CHCl_2 \qquad (51)$$

and

$$CHCl_2CFCl_2 + F\cdot \rightarrow HF + \cdot CCl_2CFCl_2 \qquad (52)$$

$$\cdot CCl_2CFCl_2 + F_2 \rightarrow F\cdot + CFCl_2CFCl_2 \qquad (53)$$

The other reactions studied by Miller may in most cases be explained without difficulty in a similar manner.

Theoretical speculations like those outlined in the greater portion of this section undoubtedly possess numerous shortcomings. On the one hand, there are many things, such as the properties of the aromatic nucleus, which an electronic formula cannot fully portray; while on the other, it has been alleged, not without point, that with a little mental dexterity such mechanisms can be made to account for every observed reaction, and for many more which are not observed. All this has its roots in the fact that the symbols we use only very roughly represent the entities for which they are supposed to stand. However, as has been said before, these theories, when labelled as such, are well worth while provided they first correlate and explain, next stimulate thought, argument, and disagreement, and finally lead to experiment and progress.

IV. FLUORINE GENERATORS

Since elementary fluorine has not yet been made available commercially[2] and in the past has almost invariably been produced as used, it would appear to be desirable to introduce here a discussion of fluorine generators. However, this field is in itself so extensive that no more than a representative survey can reasonably be attempted.

Fluorine has almost always been prepared by the electrolysis of alkali metal fluorides, mostly potassium bifluoride, KHF_2, either molten or dissolved in anhydrous hydrogen fluoride at various concentrations. Since Moissan's time about a score of generators have been described, operating under widely differing conditions and yielding, as might be expected, quite different results. While in general the procedure is not really difficult to carry out, it is safe to say that in each individual case notable obstacles have been encountered, and that the ideal apparatus probably has yet to be described. For our present purpose, five generators, representative of the more important types, will be considered in some detail here. An effort will be made to point out the respective advantages and limitations of the different units, the problems which have in general been encountered, and, in the writer's opinion, what would seem to be the direction of progress.

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[2] Note, however, the statement by Wartenberg (90a) regarding the loan to him for experimental purposes of a cylinder of fluorine (75 per cent fluorine, 17 per cent nitrogen, 8 per cent oxygen by volume) from the I. G. Farbenindustrie A.-G.

A. The Moissan generator

The first successful fluorine generator was, of course, the one described originally by Moissan (62, 63). He isolated the element for the first time by the electrolysis of an approximately 30 per cent solution of potassium bifluoride in anhydrous hydrogen fluoride maintained at a temperature between −23° and −50°C. in order to prevent volatilization of the solvent. The apparatus, which was made entirely of platinum, consisted essentially of a U-tube with side-arm delivery tubes, equipped with tight fluorite stoppers carrying the electrodes, and surrounded by a bath of liquid methyl chloride. The anode, at which the fluorine was liberated, was made of a highly resistant platinum-iridium alloy, but even this was seriously corroded. The anhydrous hydrogen fluoride required for the electrolyte was prepared by heating pure potassium bifluoride in a platinum retort, the salt having previously been desiccated with great care under reduced pressure. Every precaution seems to have been taken to exclude air and moisture during this operation and also the subsequent electrolysis. The fluorine was evolved at the rate of 1.5–2.0 liters per hour, using a current of 3.5–4.5 amp. at a potential drop of about 40 volts, but data regarding the current efficiency of the cell do not appear to have been given.

There would seem to be little doubt that the anode gas consisted of elementary fluorine in a relatively high state of purity, since it readily ignited crystalline silicon and, after being freed from entrained hydrogen fluoride, was absorbed completely by metallic mercury in the cold, as well as by heated iron contained in a platinum tube. In the latter case, the increase in weight of the iron presumed to be due to fluorine was shown to be closely equivalent to that of the hydrogen liberated at the cathode during the same period. The mercury-absorption test would seem to exclude the possibility that the anode gas contained oxygen, and perhaps oxygen fluoride as well; but the reaction with heated iron would certainly not do so, since these impurities would be absorbed readily at the same time. Had this occurred, then the fluorine actually produced would have been definitely less than equivalent to the corresponding cathode hydrogen, a result which would naturally be anticipated on account of the corrosion of the anode during the electrolysis. Many years later Ruff (72) studied the Moissan generator intensively, and found that under the best conditions the electrical efficiency with respect to fluorine was only 30 per cent, while that with respect to hydrogen was only 40 per cent, thus lending support to Moissan's observation that they were not far from the same. These very low values were found to be due to (*1*) the reaction of fluorine with the platinum anode, (*2*) the reaction of hydrogen with platinum compounds progressively dissolved by the bath, (*3*) the simultaneous liberation of some hydrogen and some fluorine in the anode and cathode chambers, respectively, and (*4*) the presence of water in the electrolyte. All of these considerations taken together leave a slight uncertainty as to the exact quantitative composition of the fluorine gas originally obtained when the element was discovered.

Some years later, Moissan (67) described a similar generator using a copper

U-tube, and giving approximately the same results more economically. He found, however, that a copper anode could not be used successfully, since a film of metal halide caused polarization.

Moissan's pioneer generator stands as the classical example of the U- or V-type cells, and of the use of a metallic anode in the production of fluorine by electrolysis. It also operated at very low temperatures with a relatively dilute but presumably anhydrous electrolyte. It had the fundamental asset of being highly successful for the purpose for which it was intended, but also possessed a number of significant disadvantages. The initial cost was very high, the expensive anode was attacked by the fluorine, the procedure, including the preparation of the anhydrous hydrogen fluoride, required highly skilled operators, and the apparatus as well as the process as described could not easily be adapted to production on a larger scale without much modification. In the decades immediately following, these practical obstacles operated to deter scientists from preparing and studying the new element, in spite of the excellence of the original work.

B. The Mathers generator

In 1919 Argo, Mathers, Humiston, and Anderson (3) described a new generator which, at the outset, gave promise of obviating most of the difficulties encountered by Moissan and has since become generally known as the Mathers generator. In this case, the fluorine was produced by the electrolysis of molten potassium bifluoride, at temperatures from 250–300°C. The apparatus consisted essentially of an electrically heated cylindrical vessel, made of heavy copper and open to the air, which contained the electrolyte and also served as the cathode. The anode was made of a thick graphite rod, suspended in the melt by means of a copper bolt, and completely surrounded by a cylindrical copper diaphragm, closed at the top, and provided with an exit tube for the fluorine as well as a series of slots near the bottom to facilitate the electrolysis. The rod which carried the anode passed directly through an insulating stuffing box filled with powdered fluorspar, which was supported on a fluorite ring and compressed by a jam nut. The diaphragm was also equipped with a removable plate at the bottom, in order to prevent any hydrogen rising from below from entering the anode chamber and reacting explosively with the fluorine.

The electrolyte was prepared by treating potassium fluoride with an excess of commercial hydrofluoric acid, and always contained water at the start. The mixture was slowly heated in an open metal container for several hours, until the temperature reached 225°C., to expel the excess of water and acid. Then the melt was transferred to the cell, where it was electrolyzed for several additional hours at a low current density in order to remove the remainder of the water; it was not considered to be very hygroscopic after being dehydrated in this manner. This preliminary electrolysis was necessary in order to avoid almost complete polarization of the cell due to the formation of an oxygen gas film on the surface of the anode. Afterwards, at a temperature of about 240°C., fluorine was evolved freely and the generator could be operated continuously at a much higher current density without polarization. The exhausted bath could not, according to the published report, be regenerated by passing hydrogen fluoride into it. Instead, it was poured out, frozen, ground, and redissolved in hydrofluoric acid, after which the heating and dehydration were repeated. In operation under favorable conditions at 225–250°C., using a current of 10 amp. at a potential drop of 12–15 volts, fluorine was evolved at a current efficiency stated to be about 70 per cent. No comment seems to have been made regarding the analytical composition of the anode gas, beyond the general implication that it consisted of nearly pure fluorine.

The writer and his students used a generator (7) of the Mathers type for a period of several years, and so can speak from experience regarding it. The advantages of this unit were clearly obvious even to the more casual observer. It was rugged in construction and relatively inexpensive, the anode was cheap and but little attacked, the procedure did not require the services of very highly skilled operators, and the apparatus could easily be adapted to larger-scale production. Unfortunately, however, this type of generator has a number of serious limitations, some of which, at least, were probably not fully realized by the original investigators, who were engaged in an effort to produce toxic gases under the pressure of wartime conditions.

Most important of all, much recent experience has made it clear that (*a*) the electrolysis of the bath at low current density, using less than 10 amp. even for the entire life of the charge, would not dehydrate it completely; (*b*) the electrolyte was markedly hygroscopic, especially on standing exposed to the air; and (*c*) unless the melt was completely anhydrous, the anode gas contained large and widely varying quantities of oxygen and oxygen fluoride. This remained true to a considerable extent even when the melt was sufficiently dry to undergo electrolysis smoothly and without polarization. These facts have been well substantiated by numerous analyses of the anode gas prepared in this Laboratory by Day (24) and by Miller (59), who has also collected and discussed a considerable body of evidence from the literature in regard to the contamination of fluorine obtained in this manner. It has even been found that high atmospheric humidity had an adverse effect upon the composition of the anode gas. While it is probable that a further electrolysis of this kind of melt at higher current densities, using more than 10 amp., would ultimately dehydrate it before the exhaustion of the charge, it certainly would not remain anhydrous for any considerable time if left standing exposed to the air, either heated or frozen. Incidentally, it has been found entirely practicable to regenerate an exhausted charge by passing anhydrous hydrogen fluoride into it, contrary to the earlier report (3). It should also be clear that the limitations just outlined would apply not only to the Mathers type of generator, but to all units in which strictly anhydrous conditions cannot be or are not maintained.

This type of generator also possesses several lesser, but still very real, drawbacks. In the writer's experience, on account of the high temperature a considerable amount of copper was dissolved, transported, deposited, or precipitated. This was found to result in much corrosion of the diaphragm, with attendant contamination of the electrolyte and the accumulation of sludge. During the operation of the cell, detonations, sometimes powerful, occurred from time to time in spite of much care. They were probably due to the recombination of hydrogen and fluorine, and usually took place in the cathode compartment, when fluorine bubbled into it through the diaphragm, presumably owing to the obstruction of the exit tube by frothing viscous electrolyte. Under these conditions the explosions were essentially harmless, but at rare intervals and for reasons not well understood, they took place in the anode compartment; this resulted in the forcible expulsion of molten electrolyte from the cell and in this case constituted a real hazard. Perhaps it should also be mentioned that the fluorite ring was very fragile, salt bridges in the anode chamber at times caused annoying short circuits, and recurring polarization was frequently time-consuming.

There can be no doubt that the Mathers generator represents a real advance along the difficult way toward fluorine production. Nevertheless, from the preceding outline it will be clear that neither this unit, nor similar ones which have followed it, when operated intermittently, open to the free access of air and moisture as described, can be relied upon to produce consistently either analytically pure fluorine or an anode gas of constant composition.

C. The Fredenhagen generator

About a decade later, in 1929, Fredenhagen and Krefft (34) described another generator, identical in principle with the Mathers unit but differing from it in several important respects, in that moisture was rigorously excluded, and an anhydrous electrolyte employed, which was repeatedly regenerated by means of anhydrous hydrogen fluoride. The electrically heated, cylindrical, copper electrolysis vessel serving as the cathode was provided with a thick cover of the same metal, which fitted tightly into a top flange by means of a polished and tapered joint. The insulated graphite anode, supported on a nickel rod, was surrounded by the usual perforated copper diaphragm, also insulated, and equipped with a hydrogen baffle at the bottom. The nature of the insulating material was not given.

The potassium bifluoride electrolyte was presumably dehydrated in the apparatus itself by a patented process (106), which consisted essentially in passing dry air through the crystalline solid salt at elevated temperatures and reduced pressure for such time that when the bifluoride was subsequently heated to a high temperature, the hydrogen fluoride evolved showed a minimum electrical conductivity. This procedure would be expected to yield a very dry charge, which was regenerated without difficulty after electrolysis by bubbling anhydrous hydrogen fluoride, made by heating dry potassium bifluoride (32), directly through it. The electrolysis was conducted over a slowly rising temperature range from 160–250°C., corresponding to a progressive change in the composition of the electrolyte by the loss of hydrogen and fluorine from $KF \cdot 1.6HF$ to $KF \cdot 0.8HF$, when it was regenerated. The fluorine was given off smoothly, using an average current of 20 amp. at a potential drop of 8.5–9 volts, at the rate of 7.6 liters per hour and a current efficiency of 92–94 per cent. The product was considered to be free from oxygen and other contaminants, but no analyses were given, and the method by which the very high efficiency was measured was not indicated.

These investigators have found that when the proportion of hydrogen fluoride in the melt was greater than that corresponding to $KF \cdot 1.8HF$, the graphite began to be wet by the electrolyte and soon disintegrated mechanically, whereas if the proportion was less than that corresponding to $KF \cdot 0.8HF$, the internal resistance of the bath was greatly increased. In the intermediate working range, however, the graphite was neither attacked nor wet by the electrolyte, and showed a constant overvoltage of 5 volts, presumably due to the presence of a very thin film of vapor at the anode surface. They also reported that the melt absorbed water greedily from the air even at 250°C., and they made some interesting observations concerning the effect of water on the process. When the melt contained more than 1 per cent, or less than about 0.3 per cent, of water, the electrolysis proceeded smoothly, yielding oxygen in the former case and essentially fluorine in the latter. In the intermediate range however, erratic and troublesome polarization occurred, due to the overspreading of the graphite by an insulating oxygen gas film. This led to potential drops up to 50 volts, accompanied by sparking through the gas film and injury to the electrode. These observations emphasize the great importance of and even necessity for the use of a strictly anhydrous electrolyte.

The Fredenhagen generator and procedure would seem to avoid the major difficulties involved in the Mathers method, since an essentially dry electrolyte was maintained anhydrous in a closed apparatus and regenerated by dry hydrogen fluoride. Unfortunately, however, no analyses of the anode gas were presented to support the claim that it was practically pure fluorine, presumably after the removal of hydrogen fluoride. The product may have been free from oxygen compounds, but undoubtedly it contained a small but unknown amount of carbon fluoride, according to earlier reports. No definite comment was made concerning anode or cathode chamber explosions, salt bridges or entrainment, and the disintegration of insulating material. It would be interesting to know if any of these difficulties were actually encountered, and if so, how they were met. These investigators have made a distinct contribution by setting forth clearly the serious and varied effects of progressively changing proportions of water in the electrolyte, and have supported these statements with a considerable body of experimental data. In addition, they have reported the range of hydrogen fluoride concentration in the melt in which a graphite anode could practicably be used. Altogether the methods and results as presented are satisfying, and in considerable measure convincing.

D. *The Cady generator*

The next generator to be considered was described recently by Cady (18), and was essentially different from the others in that it operated at a medium temperature of 75°C. and had a nickel anode. No particular effort was made in this case to exclude moisture rigorously during the process. The gas-heated, cylindrical reaction vessel, serving as the cathode, was made of Monel metal, provided with a flat flange at its upper edge and covered by a sheet-copper lid. A cylindrical copper diaphragm, closed at the top, which also served as anode support, electrical conductor, and hydrogen fluoride bubbler, was suspended from the lid by a copper tube, which passed through an insulating stuffing box filled with Portland cement. This tube was used to conduct the fluorine from the apparatus, and was also connected to the positive pole of the D.C. circuit. Four large square holes were cut in the diaphragm near its lower end, which itself was closed by a copper plate carrying the anode. The latter was made from a roll of thin sheet nickel, which was bolted to the inside surface of the plate and projected for a short distance upwards.

The electrolyte was prepared by melting a charge of technical potassium bifluoride in the cathode receptacle, at a temperature somewhat below 250°C., and subjecting the melt to a preliminary electrolysis, using a graphite anode and a current of 12–15 amp., until such time as fluorine was liberated. This required not more than 3 hr., during which time the polarization and arcing at the anode were periodically relieved by opening the circuit. Then anhydrous hydrogen fluoride was passed directly into the melt, in such a manner that the latter was always maintained at a temperature not far above its melting point, and the process was continued until the melt contained about 43.4 per cent hydrogen fluoride, corresponding fairly closely to the composition required by the formula $KF \cdot 2HF$. This procedure, which was necessary in order to avoid undue loss of hydrogen fluoride, required a progressive lowering of the temperature of the melt from 250°C. to a little above 70°C. Then, with the diaphragm in place and the lid slightly raised to avoid explosions in the cathode chamber, the cell was ready to operate. During the electrolysis, the copper diaphragm became coated with an insulating layer, and the anode gas was liberated only from the nickel. At a temperature of 73–75°C., using a current of 12 amp. at a potential drop not stated for this unit (but probably about 10 volts (19)), fluorine was produced at an average current efficiency of 72 per cent. A sample of the product contained in a glass bulb was found to be absorbed by metallic mercury to the extent of 98 per cent, but neither the conditions of sampling nor other details of this analysis were given. Interestingly enough it was observed that an electrolyzing current of 8 amp. produced sufficient heat to maintain the generator at 75°C. For lower or higher currents it was necessary to heat or cool the cell, respectively, in order to maintain the required temperature. The electrolyte was considered to be exhausted when it would no longer remain completely liquid at 75°C., and it was regenerated by bubbling anhydrous hydrogen fluoride into it through the diaphragm, until it again contained about 43.4 per cent of hydrogen fluoride. The addition was presumably controlled by titration, but this is not so stated. Cady (19) subsequently described a somewhat improved generator of the same type, which was steam heated and had an anode suspension similar to Fredenhagen's, but its performance was essentially the same as that of the apparatus just considered.

These generators, which were intentionally patterned after an older model by Lebeau and Damiens (52), were almost certainly an outgrowth of an earlier work by Cady (17), in which he presented the freezing point–composition diagram for the system KF·HF, together with related vapor pressures. In this paper he has made clear quite convincingly that, on theoretical grounds at least, there are three temperature ranges within which the electrolysis would be practicable. These lie in the vicinity of −80°, 70°, and 240°C., respectively, and represent conditions such that a considerable variation in the composition of the liquid does not affect the melting point too greatly, and also such that the corresponding vapor pressures are relatively low.

In addition to high-temperature generators, the writer has had several years' experience (8) with a medium-temperature unit using a nickel anode, which was very similar to Cady's in principle but differed from it somewhat in detail. The advantages of this type of outfit were not far to seek. First, it operated under conditions under which the melting point of the bath was least affected by changes in its composition. Then, the medium temperature employed markedly reduced corrosion, as well as heating costs and operating difficulties. Finally, the use of the nickel anode completely avoided the troublesome polarization phenomena associated with graphite, and at the same time made certain that the product could not be contaminated with volatile carbon fluorides.

On the face of it, such advantages would appear to be compelling, but unfortunately this type of cell too has inherent limitations which must be both recognized and successfully met, if satisfactory performance is to be expected. First of all, and most important, the usual smooth operation of this kind of generator, in the complete absence of polarization effects, gives no outward indication whatever of the actual composition of the anode gas, which can vary over a wide range. Without adequate analytical control, this fact alone has frequently led to disappointment and loss. We have found (8), for example, that when a generator of this general type, operating in routine fashion for many hours at 5 amp., was taken from the line and tested, the anode gas after the removal of hydrogen fluoride was found to contain only 75.5 per cent of fluorine by analysis, using the apparatus to be described in the following section. Also, a special test generator operated for a long period failed to yield good fluorine at all when a current of 5 amp. was used, but with 10 amp. or more, the melt was dehydrated in 2–3 hr., after which the anode gas contained essentially 100 per cent fluorine by analysis. Furthermore, when the same test generator, with the melt anhydrous, was tightly closed and allowed to stand hot overnight, on resumption of the electrolysis, using 5 amp., the anode gas, after 30 min., contained only fluorine. However, if the melt was allowed to stand exposed to the air, even if only to the extent of two small portholes in the cell cover, the results were quite different. Under otherwise identical conditions, on resumption of the electrolysis after standing over one night, the anode gas contained only 83 per cent of fluorine, and after standing over 48 hr., it contained only 53 per cent of the halogen by analysis. These figures would apply, of course, not only to the units under discussion, but also in equal measure to all open-type generators.

Another difficult problem encountered with the Cady-type unit was that of corrosion of the anode, accompanied by serious contamination of the electrolyte and low current efficiency. These effects have been found to be somewhat dependent upon the temperature, which had to be maintained as low as possible. To accomplish this, however, would require a careful control at all times of the concentration of hydrogen fluoride in the electrolyte within a narrow range, since too little would cause the bath to freeze, while too much would result in serious contamination of the anode gas by hydrogen fluoride. In our experience, under somewhat different conditions, the corrosion was heavy and accompanied by current efficiencies as low as 35 per cent. Also, the progressive contamination of the electrolyte, with attendant increase in viscosity, seemed to contribute to the production of powerful anode-chamber explosions which were observed from time to time. It is recognized of course that these results of ours were not necessarily strictly comparable to those to be obtained from the exact Cady model and conditions, but they have seemed sufficiently definite and analogous to be worthy of mention.

From the accumulated evidence it is clear that, while Cady's generator can produce good fluorine, there is no certainty that it will do so, especially when operated intermittently, unless moisture is effectively excluded from the anhydrous electrolyte at all times. The only exception to this would be when the unit is in actual operation using a current of 10 amp. or more, which, according to our experience, is sufficient to dehydrate the electrolyte in an open generator. Furthermore, it appears likely that the necessary continuous control of the composition of the bath within a narrow range would be time-consuming; and finally, a serious loss of nickel and consequent contamination of the melt can hardly be avoided. These considerations, it would seem, have not received sufficient emphasis in the published reports.

E. The Miller and Bigelow generator

The last generator to be considered in detail here was described by Miller and Bigelow (60), and was designed to produce essentially pure fluorine consistently by the simple expedient of operating throughout under strictly anhydrous conditions. The fluorine was generated by the electrolysis of dry potassium bifluoride in the high-temperature range, using graphite electrodes. The apparatus consisted of a heavy, cast-nickel, U-shaped vessel, provided with polished flanges at the ends, and equipped with polished cast-nickel caps, made gas-tight by lead gaskets and held firmly in place by parallel steel plates which were tightly compressed by steel bolts. The two graphite electrodes were screwed to nickel rods, which passed directly through the caps, and were held rigidly suspended by means of a resistant insulating packing made from powdered fluorspar and Portland cement. The caps were also provided with drilled holes, equipped with adequate couplings, for the escape of the fluorine and hydrogen, respectively, and a thermometer well extended directly into the main casting. The entire unit was contained in a rectangular box-shaped electric heater, so arranged that the caps were kept hot enough to melt any of the electrolyte which was entrained. The pure crystalline bifluoride was completely dehydrated in a tightly closed, vertical metal tower, which was electrically heated and provided with a thermometer well at the top and a large spout at the bottom. At first dry air (106) was passed through the charge at 150°C. under reduced pressure, and this was later displaced by fluorine gas (41) to complete the dehydration effectively, when fluorine of the highest quality was desired. Then the electrolyte was melted by raising the temperature, and finally run rapidly through the spout into the generator with the minimum possible exposure to air and moisture. In operation at 250–300°C., using a current of 5 amp. at a potential drop of 18–20 volts, this unit functioned very smoothly and produced fluorine which was 94.4–99.0 per cent (±0.1 per cent) pure by analysis, at the rate of about 2 liters per hour and a current efficiency of 61–89 per cent (±2.5 per cent), depending on the age of the charge and the efficacy of the drying. The apparatus would run about 25 hr. before the charge was exhausted and had to be replaced. Later on it was found that the electrolyte could easily be regenerated with anhydrous hydrogen fluoride, even at high temperatures, and thus it was possible to continue up to nearly 100 hr. before recharging and reconditioning finally became necessary.

The most important advantage of this generator was that it eliminated completely the contamination of the anode gas by unknown and widely variable proportions of oxygen and oxygen fluoride, and yielded consistently reproducible results, as stated above. In addition it was rugged, could be adapted reasonably well to large-scale production, was almost unaffected by corrosion on account of a protective surface coating on the nickel, and operated for long periods of time with great smoothness and no trace of violent reactions of any kind. The electrolyte was scarcely contaminated by nickel salts, and neither electrode was seriously attacked after many hours in service.

It would be a pleasure to record that this particular generator had no significant limitations, but, unfortunately, as in all the other cases discussed, such was not the fact. Perhaps the greatest difficulty encountered in its use was the tendency of less skilled operators to cut short the rather tedious dehydration process, which could not be completed within the cell itself because the maximum current capacity of about 10 amp. was too low. Under these conditions all the polarization phenomena so well described by Fredenhagen and Krefft (34) came into play, resulting in serious loss of time and possible disintegration of the electrodes, not to mention an impure product. Other difficulties were the infiltration of the electrolyte between the nickel and the graphite, causing the electrodes to crack at times, and the formation of salt bridges over the insulating packing, resulting in the partial or complete short circuit of the electrolyte and the rapid corrosion of the nickel rods at these points. Also, the actual capacity of the unit was low, while the manipulation of a relatively heavy metal unit at high temperatures was noticeably laborious. Finally, it was possible for an obstruction in the fluorine line to cause the eventual clogging of both exit tubes from the generator with frozen melt, resulting in the building up of high pressures within the tightly closed unit. This real hazard was completely overcome, however, by the use of a hydrogen flowmeter and indicator at all times. Altogether, the closed unit can be depended upon to produce fluorine of very high purity consistently and efficiently provided it is skillfully operated, but it now appears possible that the necessary maintainance of a strictly anhydrous electrolyte may well be accomplished in a somewhat simpler fashion.

F. Generalizations and tabulation of generators

In concluding this section it is hoped that the foregoing somewhat detailed description and discussion of five typical units may be considered sufficiently representative to serve as a basis for such generalizations as may be made concerning laboratory fluorine generators. There are five main topics to be considered: namely, (*a*) type of unit, (*b*) operating temperature range, (*c*) anode material, (*d*) condition and regeneration of the electrolyte, and (*e*) miscellaneous items.

In the light of our present experience, of these topics only *d* may be considered to be finally settled. There is no longer any doubt that while the evolution of fluorine is proceeding, either (*1*) the electrolyte must be completely anhydrous or (*2*) the anode gas will contain presumably considerable, and certainly unpredictable, quantities of oxygen and oxygen fluoride which are difficult to remove. The bath may be rendered anhydrous by a preliminary electrolysis for several hours when using a current of more than 10 amp. but can be maintained so *only* by using a tightly closed unit. The sole exception to this is when the respective exit gases are issuing from the apparatus at rates sufficient to counterbalance completely the diffusion of moisture in the opposite direction. The regeneration of the melt by passing commercial anhydrous hydrogen fluoride directly into it is now a universally accepted procedure, but care must be taken that the gas is really anhydrous, or much trouble may result. Also, the quantity added should be controlled carefully, either (*1*) by the addition of known amounts (weighed or measured) or (*2*) by the semiquantitative testing of the electrolyte itself.

With regard to the other topics, where final decisions cannot yet be made, it becomes necessary to apply individual judgment. It is more than probable that most of the generators which have been described can be made to yield good fluorine, provided the intelligent operator is sufficiently convinced of the value of his outfit to make the necessary adjustments, compensations, and effort. Without these, however, the results might well be very different.

Of all the different types of units which have been used, it seems likely that the diaphragm type, employing an electrically heated cylindrical container, is the most adequate, since it is compact, easy to handle, and has the greatest relative capacity for the electrolyte. It may be made of copper, nickel, Monel metal, or even steel (18), but it must be tightly closed, either by clamps with a safety tube, as suggested by Bockemüller (10), or even better by polished tapered joints, according to Fredenhagen (34). The container should not act as the cathode (10), so that the diaphragm may be open at the bottom and perhaps abbreviated in length without incurring the risk of mixing hydrogen with fluorine. In this case, however, the area of the insulated metal sheet, or perhaps cylinder, serving as the cathode must be sufficiently great to avoid the occurrence of annoying overvoltages (19) at this point. Finally, in any larger unit metal condensers should be provided, and arranged so as to return volatilized hydrogen fluoride to the melt. Such installations might also be used to maintain an atmosphere of hydrogen fluoride above the melt, and so to effect a continuous regeneration to compensate for the losses by electrolysis.

The operating temperature range of the cell and the anode material employed are definitely interrelated. The only substances thus far suggested which seem sufficiently inert to fluorine, and also inexpensive enough to be practicable for use as anodes, are nickel and graphite. The former cannot be used at the high-temperature range on account of extreme corrosion, while the latter may not be employed at the lower temperatures, since it is first wet and then physically disintegrated by this kind of bath, which often contains more hydrogen fluoride than the limiting value, already referred to, corresponding to the formula $KF \cdot 1.8HF$. A graphite anode operates smoothly in an anhydrous electrolyte, is scarcely affected by corrosion, cannot contaminate the melt with metal salts, and leads to relatively high current efficiencies, often 80–90 per cent, but it polarizes strongly when the melt contains about 0.3–1.0 per cent water (11, 34) (which may serve as a valuable signal), it may disintegrate physically, and it is somewhat difficult to suspend and to contact efficiently. By comparison, a nickel anode operates smoothly at all times in a medium-temperature bath (thus providing no signal), cannot produce any carbon fluorides, does not disintegrate physically, and is easy to contact and suspend. However, it is seriously corroded under optimum conditions and rapidly otherwise, it contaminates the electrolyte progressively with nickel salts which increase its viscosity, thus promoting explosions and requiring frequent replacements, and it leads to relatively low current efficiencies, often 60–80 per cent under best conditions, and to much lower values (35 per cent or less) otherwise.

In this connection, a recent paper by Ruff (70), summarizing the results of his many years of experience with fluorine chemistry, offers much which is significant and important, although no attempt is made to decide the present question. It is pointed out that when a metal anode is used, a surface layer of fluoride is always formed, thin with nickel, thicker with platinum, but always more or less soluble in the bath, while the fluorine itself is liberated presumably by the thermal decomposition of the fluoride or perfluoride existing at the outside surface. Since the coating is progressively dissolving in the bath, the electrode is gradually consumed. On the other hand, when a graphite anode is employed in a high-temperature cell, the situation is quite different. In this case, if the temperature is much above 250°C., the fluorine penetrates below the surface, expands the crystal lattice of the graphite, and forms the true compound $(CF)_x$ (see above), which has a high resistance. Then the surface temperature rises abruptly, until the $(CF)_x$ decomposes, often with semiexplosive violence, thus exposing a fresh surface, until at last the electrode is disintegrated. According to Ruff, a graphite anode thus penetrated by fluorine is no longer sensitive to polarization by small quantities of water in the melt. However, if a fresh electrode is used and the temperature kept at 250°C. or below, the fluorine does not penetrate into the graphite lattice, but is progressively given off from a thin adsorbed surface layer. Under these conditions the electrode is not disintegrated and is very sensitive to small quantities of water in the melt. Ruff does not believe that the resulting transitory polarization or passivity of the graphite is due to a surface layer of gaseous oxygen, as usually postulated, but rather to the initial formation of $(CF)_x$, catalyzed by the oxygen present. This would seem to be a matter of opinion only, subject to future experimental confirmation.

Taking into consideration all of these factors, the writer now believes that a fluorine generator should have graphite or carbon anodes, and operate on the lower side of the high-temperature range, perhaps from 160–250°C., as recommended by Fredenhagen. Perhaps at some future time carbon rods may become available that do not disintegrate under the conditions required at the medium range, which, under such circumstances, would seem to be more desirable and convenient.

Miscellaneous topics which might be considered at this point are very numerous, so that only a few of the more pertinent ones will be mentioned. The exit tube for the fluorine should always be wide enough and suitably placed to avoid obstruction by solid entrained by the escaping gas. Also, care should be taken in construction to avoid opportunities for the formation of salt bridges across the material insulating the anode suspension. Such formations by-pass part, or even all, of the current, often without producing the appearance of a short circuit, thus reducing the efficiency of the cell and promoting corrosion. As insulating materials for the suspension, a mixture of coarsely ground fluorspar and Portland cement, or a combination of transite and asbestos board washers and rings, seems to be satisfactory, but no material has yet been suggested which is not slowly attacked by the mixture of fluorine and hydrogen fluoride impinging upon it. In addition, it seems appropriate to mention here that in the operation of any fluorine generator, constant care should be taken to make sure that organic vapors cannot diffuse back into or hydrogen accumulate in the anode compartment between runs, especially when the cell is standing at elevated temperatures. Neglect of this precaution can cause sharp anode-chamber explosions when electrolysis is resumed, accompanied by forcible expulsion of the melt and real hazard for the operator. In some rare cases, such explosions may occur as a result of certain other obscure causes difficult to explain, but a certain safety measure in any case is to flush out the anode space with dry air just before throwing the switch.

It is hoped that the foregoing discussion of the factors involved may be of assistance in the development of more practicable and efficient units.

For the convenience of the reader, a fairly complete chronological list of the various generators which have been described in the literature has been compiled in table 2, together with their most important characteristics and the leading references.

If we may now assume, for the purpose of this review at least, that a suitable quantity of fluorine is available at reasonable convenience, the final problem,

TABLE 2

*Fluorine generators which have been described in the literature**

D = diaphragm; U and V refer to the shapes of the units; main construction is of copper unless otherwise indicated

NAME	TYPE	ANODE MATERIAL	TEMPERATURE RANGE	CURRENT EFFICIENCY	REFERENCES	REMARKS
				per cent		
Moissan	U, closed	Platinum-iridium alloy	Low	30	(62) (63)	Original all-platinum unit
Meslans	D	Platinum			(69) (39)	Hollow cooled anode; bottom of container insulated
Gallo	D, closed	Platinum	Low		(38)	Small size, using platinum crucible
Argo, Mathers, Humiston, and Anderson	D, open	Graphite	High	70	(3)	
Ruff	D, closed	Platinum	Low		(72)	Efficiency study interrupted by World War I
Meyer and Sandow	D, open	Graphite	High	68–78	(57)	Graphite container; addition of $NaHF_2$ unsatisfactory
Simons	D, open	Graphite	High		(84) (85)	Inexpensive unit Anode insulated by Portland cement
Lebeau and Damiens	D	Nickel	Medium		(52)	
Bancroft and Jones	D, open	Graphite	High	30	(4)	Magnesium metal container and diaphragm, as suggested by Mathers
Fredenhagen and Krefft	D, closed	Graphite	High	92–94	(34)	
Schumb and Gamble	D, open	Graphite	High		(83)	Similar to that of Bancroft and Jones, but using Monel metal
Dennis, Veeder, and Rochow	V, closed	Graphite	High	45 (5 amp.) 76 (9.3 amp.)	(28, 27)	Insulated cathode

Krekeler	D, closed	Graphite	High	74–77	(50)	Unusual design
					(12)	Magnesium alloy container; electrodes introduced from below and insulated by frozen melt
Bockemüller	D, closed	Graphite	High	70–85	(10)	Insulated cathode
Calcott and Benning	D, closed	Nickel	Low		(97, 95)	Container and insulated cathode of Fe, Ni, Cu, Mg, or other metals or alloys
Denbigh and Whytlaw-Gray	D, open	Graphite	High	About 80	(26)	Large anode insulated by CaF_2 and sodium silicate cement
Miller and Bigelow	U, closed	Graphite	High	80	(60)	Produces 94–99 per cent fluorine by analysis
Aoyama and Kanda	V, closed	Graphite	High	75–90	(2)	Produces 97–99.8 per cent fluorine by analysis
Henne	V, open	Carbon	Medium		(42)	Yields similar to those of Dennis; cathode chamber closed only by notched cork
Cady	D, open	Nickel	Medium	72	(18, 19)	Monel metal unit; later model (19)

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* For other reviews on generators, see Lebeau (51), Bockemüller (11), Kienitz and Lange (49), and also an interesting discussion by Cady (19).

to be dealt with in the following section, is to determine quantitatively the composition, and so the purity, of the product.

V. THE QUANTITATIVE ANALYSIS OF ELEMENTARY FLUORINE

It is a strange fact that most of the investigators who have designed and described fluorine generators have told us little or nothing about the quantitative composition of the anode gas produced, yet this may vary over a wide range without giving any readily apparent indication of such change. The phrase "nearly pure," so frequently used, has practically no scientific meaning unless supported by analysis and, as the writer has found to his cost, the fluorine gas issuing from a generator does not speak for itself.

It is believed that the only strictly quantitative method for the analysis of the anode gas from fluorine generators readily available in complete detail, is that described in 1936 by Miller and Bigelow (60), although earlier efforts to analyze this gas by Moissan himself (65), Cuthbertson and Prideaux (23), and Meyer and Sandow (57), as well as later ones by Aoyama and Kanda (2), Cady (18), and Wartenberg (90a), should be noted. Our procedure took advantage of two fundamental facts: namely, that metallic mercury, when agitated, absorbs *fluorine quantitatively at room temperature*, a fact which has been known since Moissan's time; and that an undisturbed mercury surface exposed to the gas forms a completely protective coating. The apparatus, which consisted essentially of a constant-volume gas buret, is shown to scale, with the surrounding water jacket removed, in figure 3, and the operation will be described by means of a short synopsis of the more important steps in the procedure for the analysis of a sample of fluorine.

The apparatus was first connected at M to a manometer equipped with a levelling bulb, and later at C to the fluorine line by means of a tight, metal, slip joint. At first, mercury was admitted through M and G to F, and the readings on the manometer scale corresponding to the two marks at F and the one at K were observed, with H open. This made it possible to correct later pressure readings at F for a new position of the apparatus relative to the manometer after the former had been demounted and then replaced. After this the capillary above F and the space E were filled with mercury, as indicated in the figure, thus exposing a ring of the metal to the absorption chamber above. Then fluorine was passed in at C and out through A, until the protective coating was formed on the ring above E, and equilibrium completely established. Following this, mercury from B was allowed to flow through D into the exit capillary of the apparatus as shown at A, and the fluorine stream by-passed. After a few minutes the entrance capillary was similarly closed with mercury from B, disconnected at C, and the tips at A and C covered with sealing wax. This gave a sample completely sealed in the absorption chamber in equilibrium at atmospheric pressure and water-jacket temperature. Now the buret was demounted, tipped at an angle, and the mercury repeatedly swirled into contact with the fluorine until the surface of the metal remained bright, which constituted a sharp end point. Finally, the unit was replaced, the final pressure read at F, and the value corrected by means of the new reading on the manometer scale corresponding to the mark at K, now taken with G closed and H open. From the observed drop in pressure, the percentage of fluorine in the sample could be calculated directly.

In order to analyze the residual gas, if any, mercury was allowed to enter the apparatus through M, G, and E until the pressure inside became approximately atmospheric. Then the unit was connected at B to a gas-analysis apparatus of the Ambler type (1), and the air was cleared from the line by passing mercury in at M, and G, and out from the capillary below C, through D and B. After this the stopcock D was reversed, and the residual gas was led out through the capillary below A, also through D and B into the Ambler apparatus, where it was analyzed for oxygen, carbon dioxide, and inert residue by conventional methods.

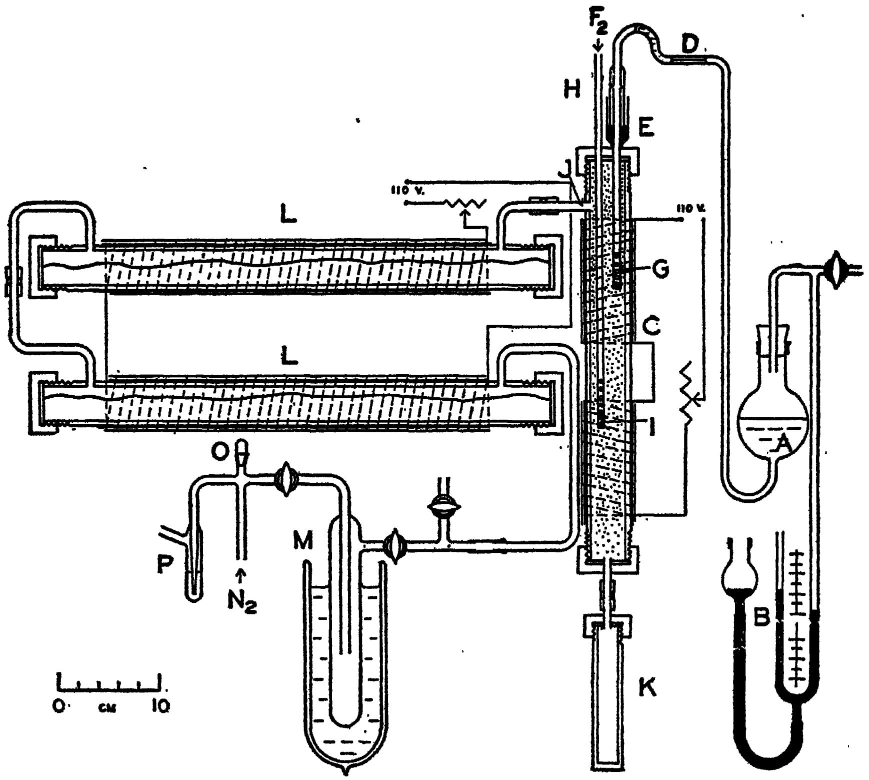

FIG. 3. Constant-volume gas buret for the quantitative estimation of elementary fluorine.

As an illustration, a series of analyses of the anode gas from our closed U-type generator (60) is shown in table 3.

The principal error involved in the fluorine analysis was that inherent in making the several manometer readings, and could have a maximum possible effect upon the final value of approximately ±0.15 per cent, but it was very likely that these errors would be at least partially compensating. Any possible errors due to pressure or temperature equilibrium at the time of sampling, or to changes in the water-jacket temperature during the analysis, or to any volume differential resulting from the formation of mercury fluoride in the buret, have been shown to be entirely negligible. Consequently it seems justifiable to consider that the final values for the volume per cent of fluorine in the sample were accurate to within ±0.1 per cent. This analytical procedure has, however, one definite limitation, which is that the sample must not contain any considerable proportion of oxygen fluoride, which would react slowly with the mercury, making the final value for fluorine only approximate. This condition is always made clearly obvious to the analyst, however, by the fact that in such cases the end point is not sharp, but drifts slowly for a considerable time. Altogether, the analysis just discussed, in the hands of a reasonably trained operator, is neither difficult nor time-consuming, and may be relied upon to give excellent results.

TABLE 3

Composition of the gas produced by the generator vs. the time of electrolysis

NO.	TIME OF ELECTROLYSIS BEFORE SAMPLING		TEMPERATURE OF ELECTROLYTE	PER CENT BY VOLUME OF			
				F_2	CO_2	O_2	Inert
	hr.	*min.*	°*C.*				
A-1	11	30	265	95.4	1.1	2.2	1.3
B-1	16	30	250	94.8	0.71	3.7	0.79
C-1	25	0	260	94.4	0.63	2.8	2.17
D-1*	7	55	249	97.1	0.97	1.6	0.43
D-2	18	50	256	98.8			
E-1	8	25	251	97.4	1.3	0.98	0.32
F-1	3	30	247	96.5	1.8	1.4	0.30
F-2	23	20	257	99.0	0.22	0.38	0.40

* The procedure for drying the charge with fluorine was introduced just before making run D-1, and the resulting improvement is clearly obvious.

VI. CONCLUSION

In the foregoing pages an attempt has been made to present (*1*) a detailed description of the various researches dealing with the action of elementary fluorine on organic compounds and free carbon which have appeared in the literature; (*2*) a summary of the theories and mechanisms which have been put forward to explain the observed results; (*3*) a discussion of the construction, operation, and characteristics of five different typical fluorine generators; and (*4*) an outline of a reliable method for the analysis of fluorine gas. In addition, pertinent data regarding compounds and generators have been tabulated for convenient reference. Throughout the presentation, an effort has been made to maintain a critical approach, and to offer some appraisal of the several researches which have been dealt with. This has been done with a clear understanding of the risks involved in passing a measure of judgment on the work of others, and with due recognition of the writer's own limitations, but it has also been done with the conviction that only in this way could the present review give full value to the thoughtful and discriminating reader.

Fluorine chemistry, including direct fluorination, has been the subject of intensive research, and has undergone great development recently, much of which has not yet been recorded. Enough is known, however, to justify the prediction that all the contributions of the last sixty years to this field, important as they are, may well appear small in comparison with the advances which will be described in the next decade or two. Indeed, should the present work seem to have fulfilled any useful purpose, it is hoped to extend it later on to cover the results of this newest research in the field.

In conclusion, looking toward the future of fluorine in organic chemistry, perhaps a bit of fascinating speculation may not be totally out of place. After all, progress begins in the mind, before it is translated into the laboratory and then the plant. In the past, inorganic chemistry was the chemistry of water solutions, but now we have systems in liquid ammonia, hydrogen sulfide, sulfur dioxide, and many others. Likewise, in the past, organic chemistry has been the chemistry of the hydrocarbons and their derivatives. Now, however, it is becoming more and more possible to envision a new organic chemistry of the fluorocarbons and their corresponding derivatives, including unsaturates, polymers, plastics, refrigerants, insecticides, and even medicinals.

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ANOMALOUS NITRATION REACTIONS

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I. INTRODUCTION

Nitration as presented in beginning textbooks of organic chemistry is an orthodox reaction characteristic of aromatic nuclei in which hydrogen is replaced by nitro groups in accordance with the rules of directive influence.

In the literature, however, there are numerous nitrations in which halogens, alkyl, alkoxyl, acyl, carboxyl, and sulfonic groups have been replaced by one or more nitro groups. In some cases, one hydrogen in an alkyl group is replaced by the nitrate radical. Some polysubstituted polyphenolic ethers are partially cleaved and oxidized by nitric acid to form substituted quinones. Occasionally, with compounds of suitable structure, cyclization takes place, along with nitration.

In general, abnormal reactions are more likely to be encountered in the nitration of polysubstituted benzene compounds—notably the polyalkylbenzenes and some of their derivatives—and polysubstituted phenolic compounds.

The formation of anomalous nitration products is influenced by such factors as the composition of the nitrating mixture,[1] the temperature of the reaction, and the number, orientation, structure, and types of groups attached to the nucleus.

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[1] Unless otherwise stated, the specifications of the acids referred to in the text are as follows: sulfuric acid, sp. gr. 1.84; nitric acid, sp. gr. 1.42; fuming nitric acid, sp. gr. 1.5.

II. MONOALKYLBENZENES

The monoalkylbenzenes yield the expected nitro compounds with the usual nitrating agents. Even *n*-octylbenzene forms *o*- and *p*-octylnitrobenzenes when nitrated at −5°C. (61). Electrolytic nitration of toluene with nitric acid in methyl alcohol solution (total acidity 30 per cent) yields some picric acid as a by-product of the reaction (3, 4).

III. DIALKYLBENZENES

The three dimethylbenzenes and the three diethylbenzenes form only the usual nitro compounds with the various nitrating agents, but with *p*-cymene and *p*-diisopropylbenzene anomalous reactions occur. There is no mention of abnormal reaction products from any *o*- or *m*-dialkylbenzenes or from the di-*n*-propylbenzenes.

The isopropyl group can be replaced with a nitro group or it can be oxidized to an acetyl group or to a carboxyl group, along with nuclear nitration. The nitration of *p*-cymene has been extensively studied in an effort to minimize side reactions and to improve the yield and purity of the 2-nitro- and 2,6-dinitro-*p*-cymenes[2] (1, 43).

Since the isopropyl group is easily oxidized, it is not surprising that methyl *p*-tolyl ketone, *p*-toluic acid, 2-nitro-*p*-toluic acid, and 2,6-dinitro-*p*-toluic acid have been reported as by-products of the nitration of *p*-cymene. *p*-Nitrotoluene and 2,4-dinitrotoluene are the principal impurities in the 2-nitro- and 2,6-dinitrocymene fractions, even when the nitration is carefully carried out at low temperatures (1, 43, 87).

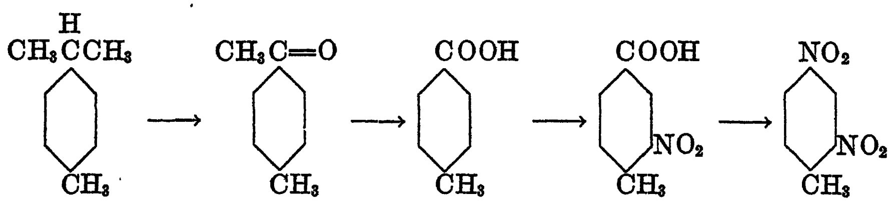
The extent of the formation of oxidation or replacement products can be controlled somewhat by varying the composition of the nitrating mixture and the temperature of the reaction. As might be expected, higher temperature favor the formation of oxidation products. At temperatures of 0°C. to −10°C. the product is mainly mono- or dinitrocymene with varying amounts of *p*-nitrotoluene and 2,4-dinitrotoluene. Kobi and Doumani (43) report 90 per cent yields of mononitrocymene containing 8 per cent of *p*-nitrotoluene by carrying out the nitration at −10°C. with a mixture of glacial acetic acid, nitric acid, and sulfuric acid.

For the preparation of 2,6-dinitro-*p*-cymene, Wheeler and Harris (87) used a mixture of fuming nitric acid and sulfuric acid below 0°C. Their product contained 24 per cent of dinitrotoluene and 6 per cent of *p*-nitrotoluene. Alfthan (1) used this same nitrating mixture at 0°C. and reported some *p*-toluic acid and some methyl *p*-tolyl ketone as by-products.

To explain the formation of both the oxidation products and the replacement products, it might appear that the following series of reactions takes place during nitration:

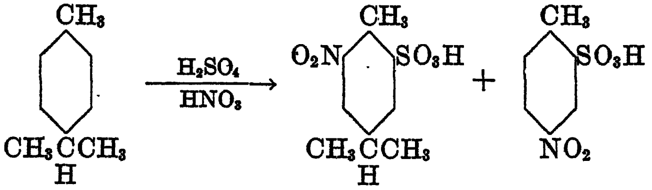
---

[2] Alfthan's paper (1) contains a survey of the literature on the nitration of *p*-cymene prior to 1920 and includes abstracts of work published in Scandinavian journals not readily available. The paper by Kobi and Doumani (43) also has a complete bibliography covering both early work and the period 1920-38.



If this mechanism is correct, either methyl *p*-tolyl ketone or *p*-toluic acid should yield some nitrotoluene or dinitrotoluene among its nitration products. Alfthan found that nitration of methyl *p*-tolyl ketone with fuming nitric acid and sulfuric acid at −10°C. yielded methyl 2-nitro-*p*-tolyl ketone, while at 60–70°C. the principal product was 2,6-dinitro-*p*-toluic acid. *p*-Toluic acid yielded either 2-nitro- or 2,6-dinitro-*p*-toluic acid but no nitrotoluenes, even under drastic conditions.

If *p*-cymene is first treated with sulfuric acid and then nitrated, the product is a mixture of the expected nitro compound and *p*-nitrotoluene-2-sulfonic acid (35):



2-Methyl-5-isopropylbenzenesulfonic acid and nitric acid also yield *p*-nitrotoluenesulfonic acid. The sulfonic group is not replaced.

Newton (47) has made a careful study of the nitration of the polyisopropylbenzenes and some of their derivatives. His nitrations were carried out with 96 per cent nitric acid in a mixture of acetic acid and acetic anhydride at 45–50°C. or with a mixture of 70 per cent nitric acid and 96 per cent sulfuric acid at 0–6°C.

In the nitration of 1,4-diisopropylbenzene, the effect of the composition of the nitrating mixture is striking, though the nitrations were carried out at different temperatures. The product from both nitrations was a mixture of the replacement product, 4-nitroisopropylbenzene, and the normal product, 2-nitro-1,4-diisopropylbenzene. With nitric acid, acetic acid, and acetic anhydride, the products were present in the mole ratio 1.44:1, while with nitric acid and sulfuric acid the ratio was 4.78:1.

Nitration of 1,4-diisopropylbenzene-2-sulfonyl chloride with a large excess of 96 per cent nitric acid at 0–5°C. yielded 4-nitroisopropylbenzene-2-sulfonyl chloride.

No abnormal products were reported from the nitration of 1,3-diisopropylbenzene or any of its derivatives.

Nitration of the chlorocymenes and the chlorocumenes has resulted in the replacement of the isopropyl group by a nitro group. Dinitro-2-chlorotoluene is an impurity in the preparation of 5,6-dinitro-2-chlorocymene with 25 per cent fuming sulfuric acid and fuming nitric acid at 0°C. as the nitrating agent (45).

The nitration of 2,5,6-trichloro-*p*-cymene with fuming nitric acid at 100°C. yields some 2,5,6-trichloro-3,4-dinitrotoluene (53). The 2,3,5,6-tetrachlorocymene yielded 2,3,5,6-tetrachloro-4-nitrotoluene under the same conditions or with a mixture of nitric and sulfuric acids. The isopropyl group is also replaced in pentachlorocumene, 2,4,5,6-tetrachlorocumene, and 2,4,5-trichlorocumene.

IV. TRIALKYLBENZENES

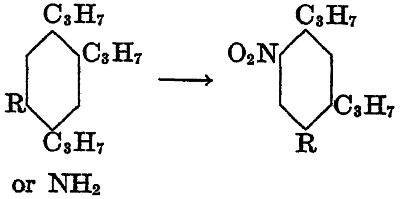
Only the usual nitration products have been reported from the trimethylbenzenes, the triethylbenzenes, and the triisopropylbenzenes, but with trialkylbenzenes containing tertiary radicals, replacement reactions occur. Unfortunately, the structures of some of the hydrocarbons were not definitely established and often neither the composition of the nitrating mixture nor the experimental conditions were recorded.

Isopropyl-4-*tert*-butyltoluene (prepared by the reaction of 4-*tert*-butyltoluene with isopropyl chloride and aluminum chloride or with isopropyl alcohol and sulfuric acid) loses an isopropyl group when nitrated with fuming nitric acid below 0°C., but isopropyl-3-*tert*-butyltoluene prepared in the same way from 3-*tert*-butyltoluene yields dinitroisopropyl-3-*tert*-butyltoluene (8, 9). Other isomeric isopropyl-*tert*-butyltoluenes yielded only normal nitration products. The replacement of the isopropyl group in the one isomer only is interesting, since Newton found no replacement products with either 1,2,4- or 1,3,5-triisopropylbenzene or with *m*-diisopropylbenzene. The hydrocarbon could have been predominantly 2-isopropyl-4-*tert*-butyltoluene, provided no rearrangement took place during alkylation.

Butyl-*m*-xylene (presumably 1,3-dimethyl-5-*tert*-butylbenzene) yields some trinitro-*m*-xylene, and *tert*-amyl-*m*-xylene yields a notable amount of trinitro-*m*-xylene (22). No experimental details are given, but it is likely that the nitrating agent was nitric acid at 70°C. or nitric acid and fuming sulfuric acid at 55°C., as described in earlier papers on the nitration of some *m*-dialkylbenzenes (21).

1,3-Dimethyl-5-(triethylmethyl)benzene yields only trinitro-*m*-xylene when nitrated with 94 per cent nitric acid and 15 per cent fuming sulfuric acid at 0–70°C. (10). It can be nitrated with absolute nitric acid in acetic anhydride at 10–80°C. to form trinitro-1,3-dimethyl-5-(triethylmethyl)benzene.

An isopropyl group is replaced during the nitration of some derivatives of 1,2,4-triisopropylbenzene (47) with either a large excess of 96 per cent nitric acid or 70 per cent nitric acid in sulfuric acid.



R = SO_2Cl, CN, or NH_2

There is a noticeable difference in the ease of nitration of some of the butyl-*m*-xylenes, due perhaps in part to the orientation of the radicals and in part to the structure of the butyl groups (48). 1,3-Dimethyl-5-*tert*-butylbenzene nitrates readily to yield the solid trinitro derivative, but the nitration product of 1,3-dimethyl-4-*n*-butyl- and 4-*sec*-butylbenzenes under the same conditions is an oil which deposits a few crystals of trinitro derivative only on long standing. The oils have not been investigated.

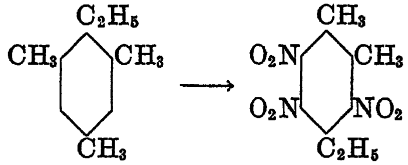
V. POLYALKYLBENZENES

Nitration of the alkylbenzenes above the trialkylbenzenes and some of their halogen derivatives can be carried out so as to yield normal nitration products in some cases, or abnormal products of two types: (*1*) those in which one or two alkyl groups have been replaced by nitro groups, or (*2*) those in which a hydrogen in one of the radicals has been replaced by $—ONO_2$ to form a benzyl nitrate.

The type of product formed by the nitration of these compounds depends upon their structure, the amount and composition of the nitrating agent, the time of reaction, and the temperature of the reaction. Nitration with fuming nitric acid alone at 0°C. to −11°C. favors the formation of nitrates, while a mixture of fuming nitric acid and sulfuric acid at 0–30°C. favors replacement products.

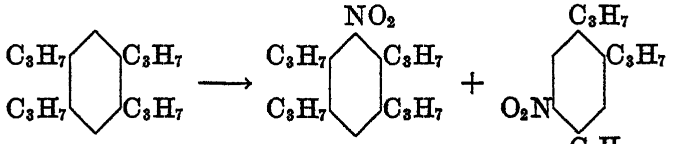
A. Replacement of alkyl groups

No irregularities have been reported among the nitration products of the three tetramethylbenzenes or the three tetraethylbenzes, but with ethylmesitylene apparently one methyl group is replaced and a methyl group migrates during nitration with fuming nitric acid and sulfuric acid (74):



Tohl and Tripke (83) reported 3,5-dinitroethylmesitylene from the reaction between this hydrocarbon and fuming nitric acid, but Smith and Kiess (74) established its identity as a trinitro compound and state that it is not trinitroethyl-*p*-xylene or trinitroethyl-*m*-xylene, but that it may be 3,5,6-trinitro-4-ethyl-*o*-xylene. These authors also report that vigorous nitration of the ethylpseudocumenes causes replacement of the ethyl group, with the formation of large amounts of trinitropseudocumene.

The effect of temperature on the yield of replacement product and normal product is striking in the nitration of 1,2,4,5-tetraisopropylbenzene with nitric acid and acetic anhydride (47). At 45°C. the yield of replacement product is 85 per cent, while at 80°C. it is the only product.

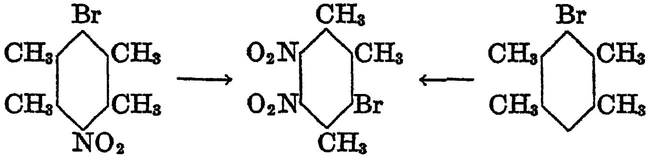


The most striking anomalous reactions occur when pentamethylbenzene, pentaethylbenzene, hexamethylbenzene, and hexaethylbenzene are nitrated. The pentamethylbenzene and hexamethylbenzene yield dinitroprehnitene (29, 72, 90), while pentaethylbenzene and hexaethylbenzene yield *p*-dinitrotetraethylbenzene (26, 40, 71). The nitrating agent can be either fuming nitric acid alone or fuming nitric acid and sulfuric acid in chloroform solution. Yields of the dinitro compounds are 22–75 per cent, with the best yields from the pentaalkylbenzenes.

Gottschalk (29) suggested that pentamethylbenzene first rearranged to prehnitene and hexamethylbenzene under the influence of the sulfuric acid, and that dinitroprehnitene was formed by nitration of the prehnitene. The principal objection to this mechanism is that pentamethylbenzene does not rearrange rapidly enough at the low temperatures and the short time for the nitration to yield an appreciable amount of prehnitene (72).

Smith and Harris (72) suggest that pentamethylbenzene may first be oxidized to prehnitenecarboxylic acid and that this acid reacts to form dinitroprehnitene with the loss of carbon dioxide. As they point out, however, prehnitenecarboxylic acid does not react with nitric acid at 0–5°C. but requires a temperature of 25°C.

One of the methyl groups in bromonitrodurene and in bromodurene is replaced by the action of fuming nitric acid to form 3-bromo-5,6-dinitropseudocumene (78):



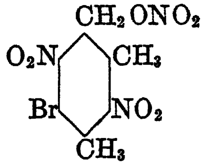
B. Nitrate formation

The possibility of the formation of nitrates during the nitration of polysubstituted benzenes was first recognized by Huender (38) in the course of an extensive study of nitrohalogenopseudocumenes. He found that 5-bromopseudocumene yielded the compound $C_9H_8O_7N_3Br$ (m. p. 150°C.) (I) when heated with a mixture of sulfuric acid and fuming nitric acid rather than 3,6-dinitro-5-bromopseudocumene, which was obtained when the nitration was carried out at ice-bath temperatures. Compound I could be obtained more readily by heating 3,6-dinitro-5-bromopseudocumene with the nitrating mixture. Huender heated I with ammonia and obtained the product $C_9H_7O_5N_2Br$ (m. p. 190–191°C.), which he did not identify. He prepared several other derivatives of I but did not identify any of them.

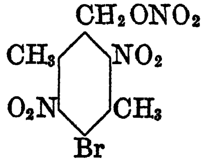
Since the analysis of I corresponded to $C_9H_8O_4N_2Br + NO_3$, or 3,6-dinitro-5-bromopseudocumene with one hydrogen replaced by $—NO_3$, Huender referred to it as a nitrate, though in conclusion he stated that it was either (*1*) a compound with the group $—CH_2NO_2$ in the side chain, (*2*) an addition product of 3,6-dinitro-5-bromopseudocumene and nitric acid, or (*3*) a nitrate.

He found that 3,6-dinitro-5-chloropseudocumene and 3,5-dinitro-6-bromopseudocumene also formed a nitrate, but that 5,6-dinitro-3-bromopseudocumene did not react with nitric acid.

Nothing further was done with these compounds until 1937, when Rinkes (60) established the structure of Huender's nitrate (I) not only as a nitrate, but as a mixture of two isomeric nitrates (II and III):



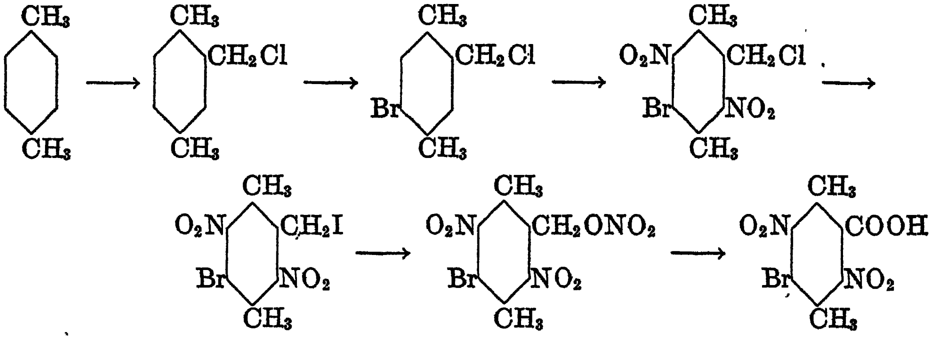
II



III

He identified Huender's product from the nitrate and ammonia as an aldehyde and oxidized it to an acid which proved to be a mixture of isomeric acids. The acids were synthesized and a mixture of equal amounts of them melted at 229°C., as did the acid from Huender's nitrate.

Rinkes synthesized the nitrate II by the following series of reactions:

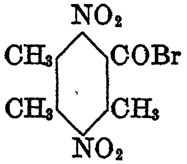


II

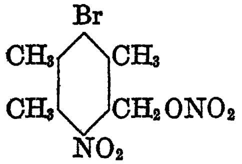
The nitrate III was synthesized in the same way from *m*-xylene. A mixture of equal amounts of the two nitrates melted at 153.5°C.

Shortly before Rinkes' work was published, Smith and coworkers had established that nitrates were formed by the action of fuming nitric acid on bromopentamethylbenzene (72) and bromodurene (78).

Smith and Tenenbaum (78) were using the directions of Willstatter and Kubli (90) for the preparation of dinitrodurylic acid bromide (IV) from bromodurene and fuming nitric acid at 0°C. They found that the compound definitely was not an acid bromide, and tentatively proposed the structure V for it.



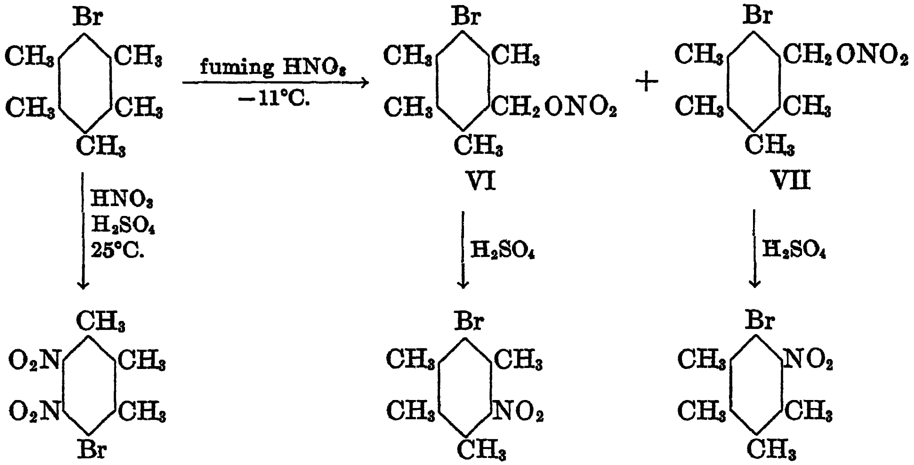
IV



V

Smith, Taylor, and Webster (77) established the structure and identity of V as a nitrate by its reactions and by synthesis.

Smith and Horner (73) obtained a mixture of nitrates from bromopentamethylbenzene with fuming nitric acid, but with fuming nitric acid and sulfuric acid the product is 4-bromo-5,6-dinitrohemimellithene.



The structures of the nitrates were established by synthesis and by their reactions, as in previous work.

Since the group —CH_2ONO_2 in VI and VII can be replaced by a nitro group by the action of sulfuric acid at room temperature, Smith and Horner suggest that nitrates are intermediates in the replacement of methyl groups by nitro groups.

Smith and Guss (71) nitrated 3,6-dibromo-1,2,4,5-tetraethylbenzene by the same procedure as for the other halogenopolyalkylbenzenes and obtained a nitrate to which they assign tentatively the structure 3,6-dibromo-2,4,5-triethylbenzyl nitrate. If this is correct, then a methyl group has been removed from an ethyl group during nitration.

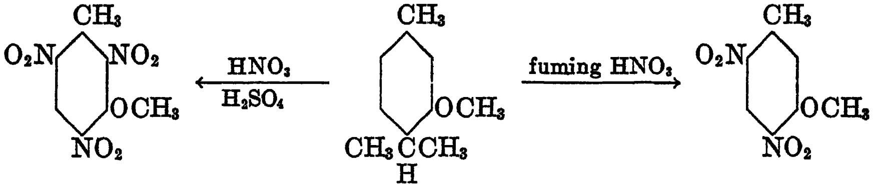
The conclusion of Smith and coworkers that the formation of nitrates is a general property of some of the highly alkylated benzenes and certain of their derivatives is amply justified. There must be some relationship between the orientation of the groups on the nucleus and nitrate formation, for not all of the isomeric bromopseudocumenes or bromotetraalkylbenzenes form nitrates.

VI. PHENOLS AND PHENOL ETHERS

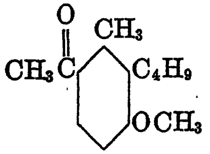
There are numerous examples of the replacement of halogens, alkyl, formyl, and acetyl groups by nitro groups during the nitration of polysubstituted phenol methyl ethers, along with demethylation and oxidation. Many of these anomalous reactions were encountered in the search for musk substitutes among the various polynitro compounds, while others were observed in connection with the preparation of nitro compounds for synthetic work or for purposes of identification.

The replacement of the isopropyl group of thymol by a nitro group (2) was observed in an attempt to prepare trinitrothymol by the nitration of 2,6-dinitrothymol with nitric acid and sulfuric acid. The product was always 2,4,6-trinitro-*m*-cresol.

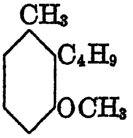
A careful study of the nitration of thymol and thymol ethers (28) proved that the isopropyl group is replaced by a nitro group even when the nitration is carried out with fuming nitric acid at low temperatures. The products were nitro-*m*-cresol ethers rather than trinitrothymol, as reported by early investigators.



Barbier (6) was interested in the synthetic musk problem and extended it to the nitration of *tert*-butylacetyl-*m*-cresol methyl ether. He prepared the ketone from *tert*-butyl-*m*-cresol methyl ether by the usual Friedel–Crafts procedure and used the structure I for it in his discussion,



I



II

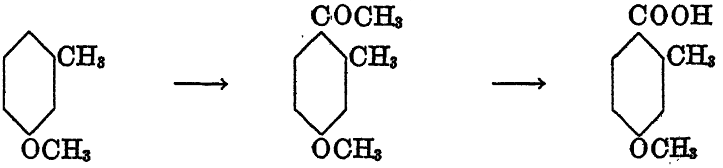
based on II as the structure for Baur's *tert*-butyl-*m*-cresol methyl ether, without proving any of the structures. As will be shown later, these structures and the structures Barbier assigned to compounds obtained from them are not correct.

The nitration of the ketone (I) with 92–95 per cent nitric acid at −10°C. yielded two dinitro compounds, expressed by Barbier as follows:

$$\underset{\text{I}}{C_6H_2\begin{cases}C_4H_9\\COCH_3\\OCH_3\\CH_3\end{cases}} \longrightarrow \underset{\substack{\text{III}\\\text{Musk ambrette}}}{C_6H\begin{cases}C_4H_9\\OCH_3\\CH_3\\(NO_2)_2\end{cases}} + \underset{\text{IV}}{C_6H\begin{cases}COCH_3\\OCH_3\\CH_3\\(NO_2)_2\end{cases}}$$

Compound III is identical with musk ambrette, obtained by Baur from the nitration of *tert*-butyl-*m*-cresol methyl ether.

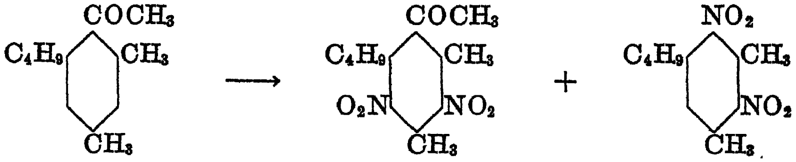
Barbier prepared acetyl-*m*-cresol methyl ether (V) from *m*-cresol methyl ether, acetyl chloride, and aluminum chloride (6).



He states that the ketone should have the structure V, since it could be oxidized to a methylmethoxybenzoic acid melting at 176°C., the melting point of 2-methyl-4-methoxybenzoic acid described in the literature. Noller and Adams (49) accept this structure for their ketone, prepared from *m*-cresol methyl ether, acetic anhydride, and aluminum chloride. This structure still may be open to question however, since Thomas (80) states that with *m*-alkyl cresol ethers and acid chlorides the acyl group goes ortho to the ether linkage. According to Kranzlein (44), phenol ethers and acid chlorides follow the rule of ortho and para substitution. He cites as an example the reaction of *m*-cresol methyl ether with bromoisobutyryl bromide, which yields a mixture of *p*- and both *o*-bromoisobutyro-*m*-cresol methyl ethers. Nitration of V yielded a dinitro compound which was identical with IV.

In a subsequent paper, Barbier lists his own and other examples in the literature in which various groups have been eliminated and replaced by a nitro group during nitration and makes the following generalization: When a highly substituted benzene nucleus is nitrated, the group meta to the first entering group is eliminated and is replaced by another nitro group (7).

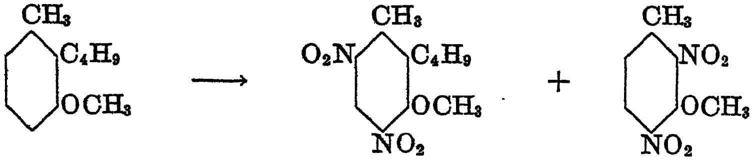
This rule holds in the nitration of 6-*tert*-butyl-2,4-dimethylacetophenone (VI), from which Barbier obtained musk ketone (VII) and 2,4-dinitro-5-*tert*-butyl-*m*-xylene (VIII), the structures of which had been established by Baur (13).



VI VII VIII

Its application to the nitration of acetylbutyl-*m*-cresol methyl ether (I) is uncertain, owing to the fact that the structure (II) which Barbier assigned to his starting material is not correct.

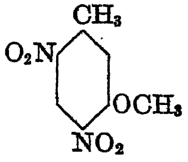
Barbier writes the nitration of *tert*-butyl-*m*-cresol methyl ether as follows (7):



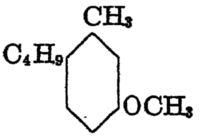
II IX X

but structures IX and X cannot be correct, since II is wrong.

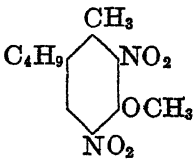
De Capeller (22) soon pointed out that X was 4,6-dinitro-*m*-cresol methyl ether (XI) rather than the 2,6-dinitro compound.



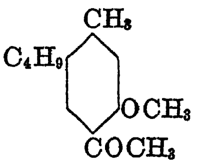
XI



XII



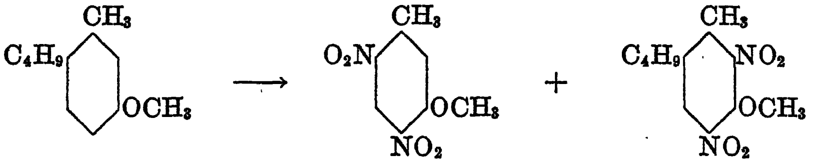
XIII



XIV

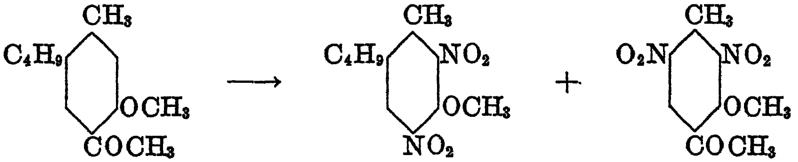
Darzens and Levy (20, 42) have established the structure of *tert*-butyl-*m*-cresol methyl ether as XII rather than II and musk ambrette as XIII, and assign XIV to the ketone I but without proof for the latter.

On this basis the nitration of XII can be written as follows:



XII XI XIII

This is in accordance with Barbier's rule. The nitration of XIV should then be written,



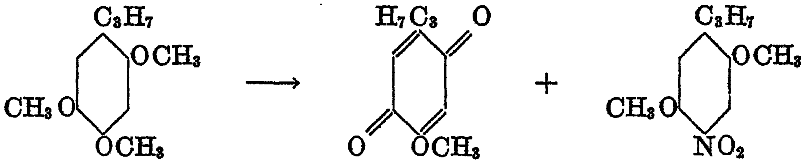
XIV XIII XV

but XV is a structural isomer of the dinitroacetyl-*m*-cresol methyl ether which Barbier obtained from the ketone V, the structure of which he and also Noller and Adams accepted as 2-methyl-4-methoxyacetophenone.

Barbier cites the formation of dinitroprehnitene from pentamethylbenzene as an exception to his rule. Another notable exception is the formation of *p*-dinitrotetraethylbenzene from pentaethylbenzene.

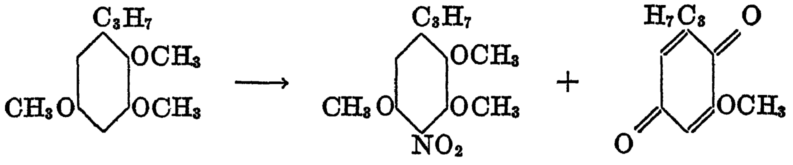
The isopropyl and *tert*-amyl radicals are replaced by nitro groups when isopropyl and *tert*-amyl-*m*-cresol methyl ethers are nitrated with nitric acid in glacial acetic acid (12).

Some interesting reactions take place during the nitration of polyphenolic ethers (81, 82, 86). When 1-*n*-propyl-2,4,5-trimethoxybenzene (XVI) reacts with fuming nitric acid at −18°C., the principal product is the quinone (XVII):



XVI XVII XVIII

and a small amount of the nitro compound (XVIII) in which one methoxyl has been replaced by a nitro group. If the reaction is carried out with 45 per cent nitric acid in glacial acetic acid at 50°C., the nitro compound (XVIII) is the main product and the quinone (XVII) is a by-product. If the isomeric 1-*n*-propyl-2,3,5-trimethoxybenzene is nitrated with 25 per cent nitric acid at 25°C., there is some demethylation and oxidation but no replacement of methoxyl groups (82).



XIX XX XXI

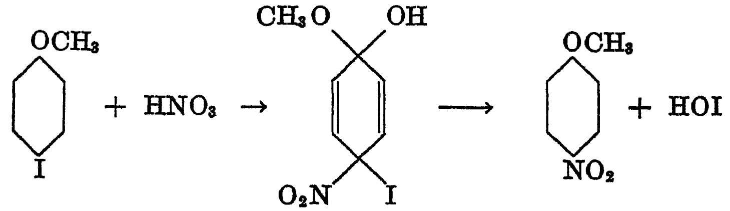
The ratio of XX to XXI is 2:1. The structures of the phenolic ethers, the nitro compounds, and the quinones were established by independent syntheses.

VII. HALOGENATED PHENOLS AND PHENOL ETHERS

The replacement of halogens by nitro groups in the halogenated phenols and phenol ethers is a common occurrence. The halogens are readily replaced when they are ortho or para to the hydroxyl or alkoxyl group, and in the order iodine, bromine, chlorine. A halogenated quinone is sometimes formed by either oxidative cleavage of the ether or oxidation of the phenol, and in some cases the halogen migrates to another position in the ring. Typical examples of these reactions are listed in table 1.

Robinson (65) suggested the following mechanism for the replacement reaction:

(1)

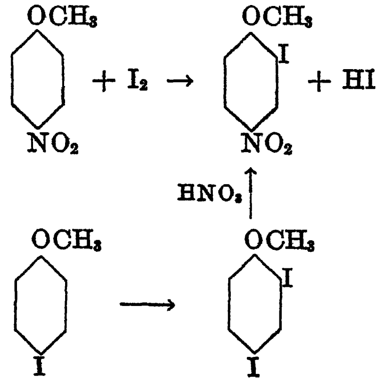


(2) $HOI \rightarrow I_2 + HOI_3$

(3) $HOI + HI \rightarrow I_2 + H_2O$

(4a)

(4b)



The iodine which separates during these nitrations will react either with *p*-nitroanisole according to equation 4a or with *p*-iodoanisole according to equation 4b.

TABLE 1

Nitration of phenols

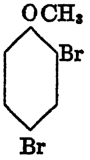
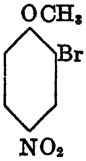
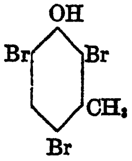
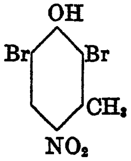
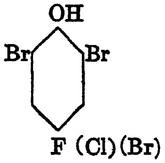
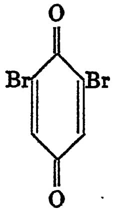
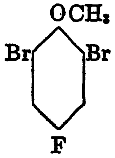
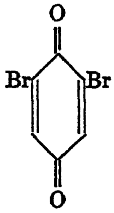
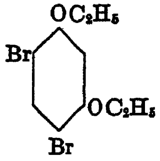
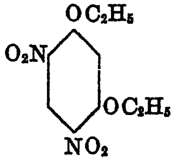
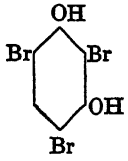
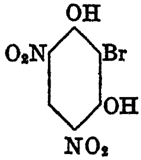
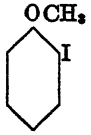
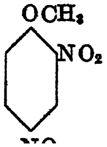
COMPOUND	REAGENT	PRODUCT	REFERENCES
	HNO_3		(15)
	HNO_3, CH_3COOH		(19, 54)
	HNO_3 (fuming)		(36)
	HNO_3 (fuming)		(36)
	HNO_3 (fuming)		(39)
	HNO_3 (fuming)		(39)
	HNO_3 (fuming)		(41)

TABLE 1—*Continued*

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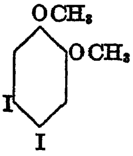
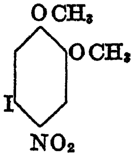
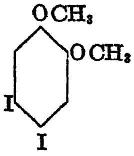
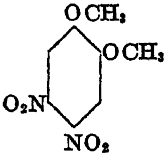
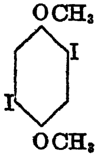
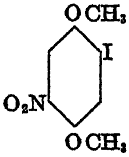
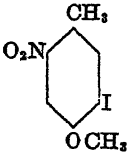
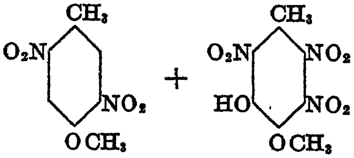
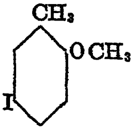
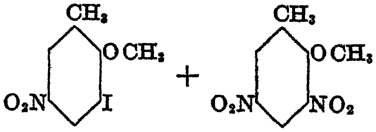
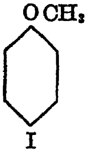
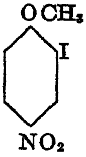
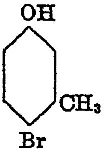
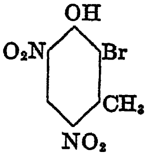
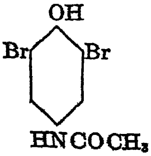
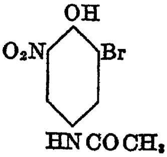
	HNO_3, CH_3COOH		(65)
	HNO_3 (fuming)		
	$HNO_3 \cdot$ CH_3COOH		(65)
	HNO_3 (fuming)		(65)
	HNO_3		(65)
	HNO_3 (fuming)		(56)
	HNO_3		(68)
	HNO_3		(63)

TABLE 1—*Concluded*

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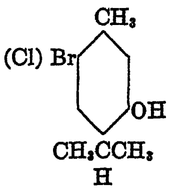
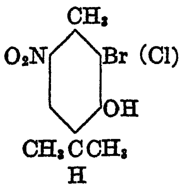
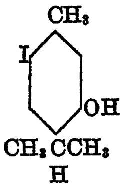
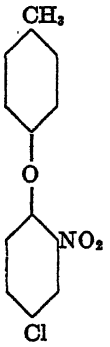
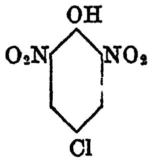
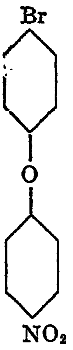
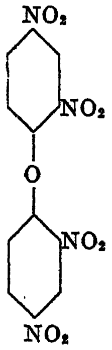
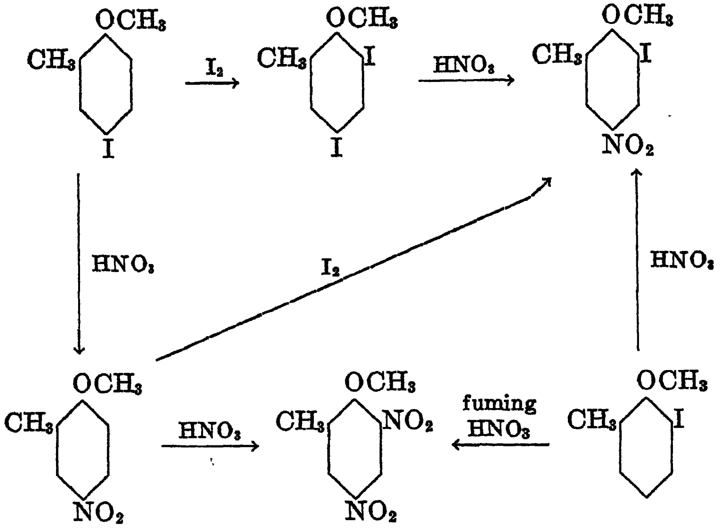
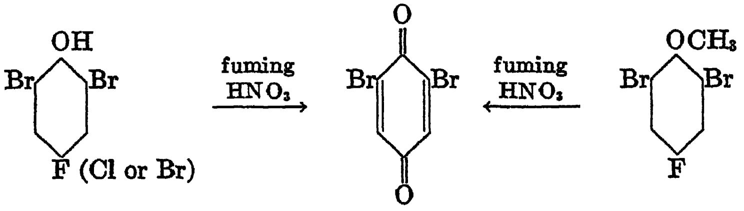
	HNO$_3$, CH$_3$COOH		(64)
	HNO$_3$, CH$_3$COOH	Mixture of nitrothymols	(64)
	HNO$_3$(fuming)		(25)
	HNO$_3$		(55)

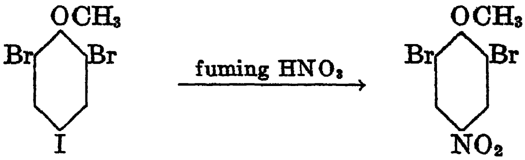
The concentration of the nitric acid appears to be rather critical for iodination to take place and, according to Robinson, the compound must have a strongly positive group, since otherwise the iodine will not be displaced.

The 5-iodo-*o*-tolylmethyl ether may react as follows:

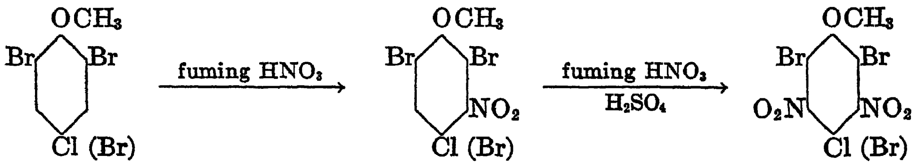


The 4-halogeno-2,6-dibromophenols and anisoles yield either halogenoquinones, normal nitration products, or, in the case of 4-iodo-2,6-dibromoanisole, a replacement product (36):





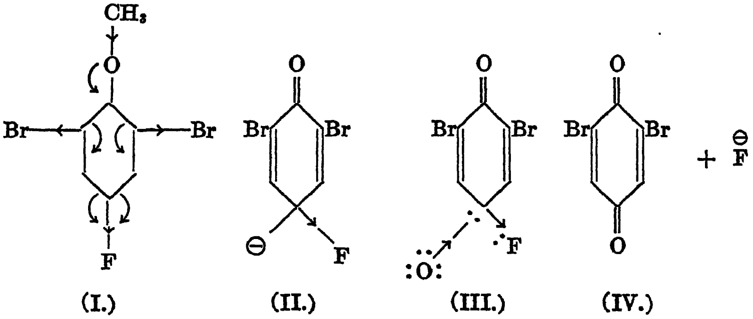
The 4-chloro- and 4-bromo-2,6-dibromoanisoles were nitrated in the 3-position. This mononitro compound on further nitration yielded a 3,5-dinitro compound:



Hodgson and Nixon (36) have explained the behavior of these compounds as follows:

> "The mechanism of reaction in the case of 4-fluoro-2:6-dibromoanisole can be regarded as a three-stage operation, *viz.*, the combined general effect . . . of the three halogens ren-

dering the methyl group sufficiently mobile (positive) for detachment and oxidation (I); rearrangement of the residue to the quinonoid form (II), facilitated by the same general effect; and subsequent attack by oxygen (III) with the separation of ionic fluorine (IV). The same mechanism will also apply to the phenols.

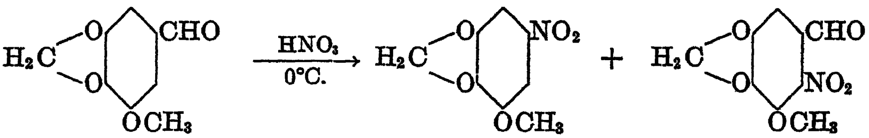


"In the 4-chloro- and 4-bromo-2:6-dibromoanisoles, the combined general effect of the halogens is inadequate to produce the necessary reactivity (positivity) of the methyl group, and so normal nitration occurs. The above general effect is still further reduced in 4-iodo-2:6-dibromoanisole, and consequently the methyl group remains intact, but the electron-donating power of the iodine is now sufficient to enable kationoid attack by the nitric acid to take place at the 4-carbon atom with nitration and elimination of iodine. Fluorine exerts a far greater general effect than the other halogens. . . ."

VIII. PHENOLIC ALDEHYDES AND KETONES

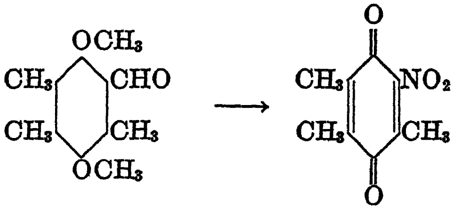
There are a few cases among the polyphenolic aldehydes and ketones in which a formyl or acyl group has been replaced by a nitro group.

An interesting example of this replacement is in the nitration of the methylenedioxybenzaldehydes (67):



Piperonal, vanillin, and anisaldehyde behave similarly (14, 23). The yield of replacement product is around 30 per cent.

The nitration of 2,5-dimethoxy-3,4,6-trimethylbenzaldehyde with nitric acid yields 3,4,6-trimethoxy-2-nitrobenzoquinone (70).



The acetyl group and the benzoyl group can be replaced by the action of nitric acid on highly substituted ketones of various types. Yields of replacement products are good in some cases. Examples are listed in table 2.

IX. AROMATIC ACIDS

Replacement products and occasionally cyclic compounds may be obtained from the nitration of aromatic acids, notably the phenolic acids and substituted acids, along with the expected nitro compounds.

A. Homologs of benzoic acid

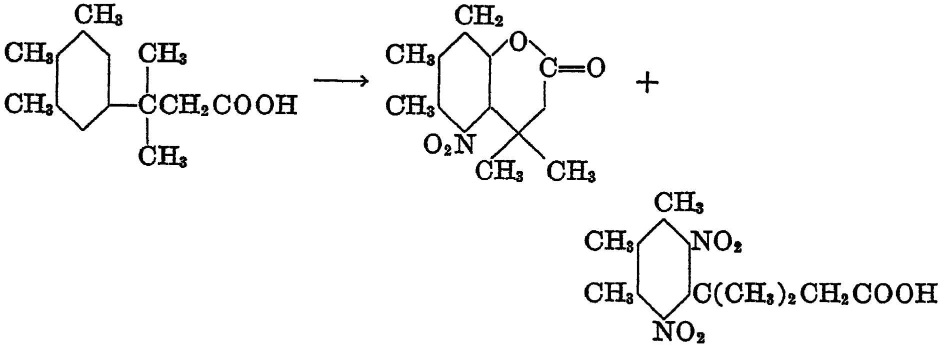
There are few abnormal products reported from the nitration of polyalkylbenzoic acids. The carboxyl group is not displaced to any extent in the polymethylbenzoic acids other than prehnitenecarboxylic acid, though Gissmann (27) remarks that durylic acid may give a small amount of nitrotrimethylbenzene when nitrated with strong nitric acid.

TABLE 2

Nitration of polysubstituted ketones

COMPOUND	REAGENT	PRODUCT	REFERENCES
2-Methyl-4,5-dimethoxyacetophenone	HNO_3, CH_3COOH	2-Nitro-4,5-dimethoxytoluene	(32)
3,4-Dimethoxyacetophenone	HNO_3, CH_3COOH	1,2-Dimethoxy-4,5-dinitrobenzene	(31)
2,6-Dibromo-4-aminobenzophenone	HNO_3	2,6-Dibromo-4-nitronitramine	(24)
2,6-Dibromo-4-aminoacetophenone	HNO_3	2,6-Dibromo-4-nitronitramine	(24)
3,4,5-Trimethoxyacetophenone	HNO_3	3,4,5-Trimethoxynitrobenzene	(31)

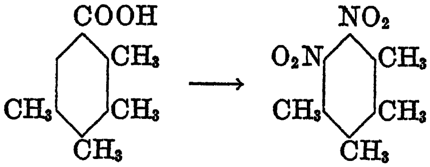
β-(3,4,5-Trimethylphenyl)isovaleric acid or its methyl ester reacts with fuming nitric acid and sulfuric acid at 10°C. to yield 4,4,6,7,8-pentamethyl-5-nitrohydrocoumarin, along with a compound whose composition agrees with that of 2,6-dinitro-β-(3,4,5-trimethylphenyl)isovaleric acid (75, 76).



Nitration of the ester with potassium nitrate and sulfuric acid also yielded the hydrocoumarin.

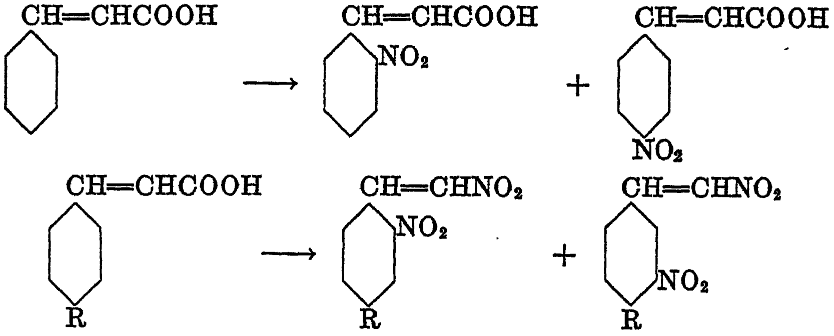
β-(3,5-Dimethylphenyl)isovaleric acid or its methyl ester yielded only the expected nuclear nitration products and no hydrocoumarin (76).

Dinitroprehnitene is formed quantitatively from prehnitenecarboxylic acid, fuming nitric acid, and sulfuric acid at 10°C. (5, 72).



B. *Cinnamic acids*

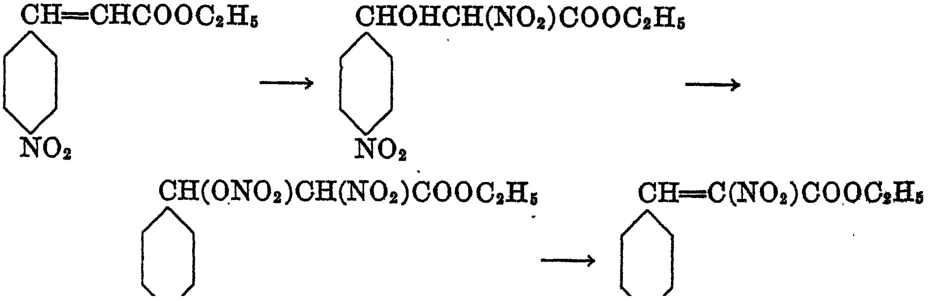
Cinnamic acid nitrates normally with nitric acid to yield a mixture of the *o*- and *p*-nitrocinnamic acids, but when absolute nitric acid (99.7 per cent HNO_3, sp. gr. 1.5204) is used at 0°C. the product is a mixture of β-2-dinitrostyrene and β-4-dinitrostyrene (84):



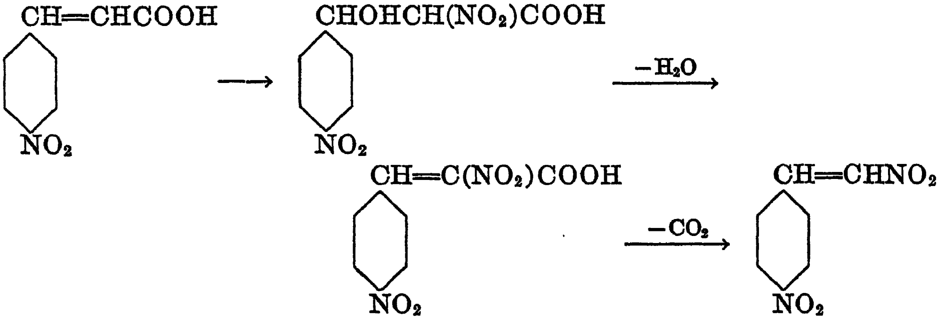
R = NO_2, Cl, Br, NH_2, OCH_3

If there is a nitro group in the 2- or 3-position, there is no further nitration of the nucleus. Even β-chloro- and β-bromocinnamic acids yielded some β-chloro-(bromo)-β-*p*-dinitrostyrene (51). Yields of the β-nitrostyrenes were 50 to 75 per cent.

Nitration of ethyl *p*-nitrocinnamate with absolute nitric acid at 0°C. proceeds as follows (84):



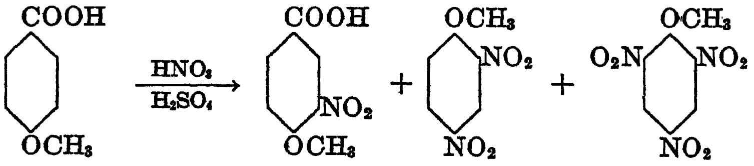
I and II were isolated and their structures established. Van der Lee suggests that the cinnamic acids react similarly:



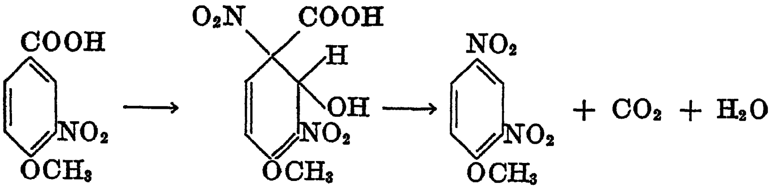
C. *Phenolic acids*

The replacement of the carboxyl group by a nitro group during the nitration of phenolic acids has been noted by many investigators (17, 23). In general, the carboxyl is more readily replaced when there are two or more hydroxyl groups or other ortho–para-directing groups on the nucleus, yet the nitration of 2,3,4-trimethoxybenzoic acid yields only the normal nitration product. The nitration of 3,4,5-trimethoxybenzoic acid leads to a mixture of the normal nitration product and 5-nitropyrogallol trimethyl ether, but if the nitration is carried out in the presence of a trace of hydrogen peroxide only the normal nitro compound is formed (30).

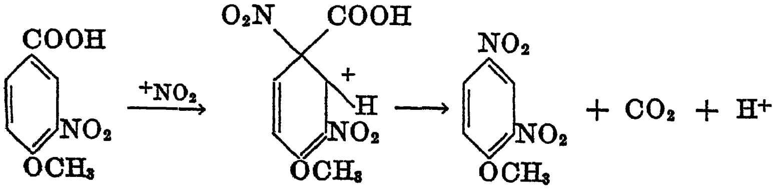
The nitration of anisic acid with sulfuric acid in absolute nitric acid yields three products (18, 23, 66):



Van der Lee (84) explained the replacement as follows:



The reaction can be explained on the basis of the cationoid mechanism of nitration (52):



Examples of the various types of replacement reactions are listed in table 3.

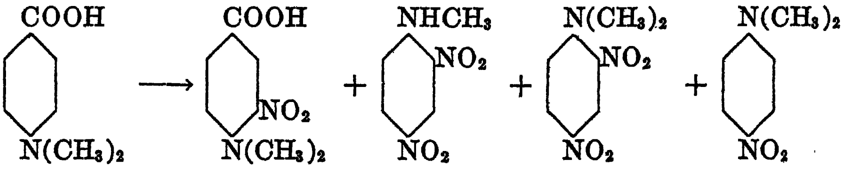
D. Aminobenzoic acids

The nitration of *p*-dimethylaminobenzoic acid can lead to a variety of products, depending on the composition of the nitrating agent and the temperature of the reaction. With a mixture of sulfuric acid and 60 per cent nitric acid at 5–10°C. the normal nitration product, 3-nitro-4-dimethylaminobenzoic acid, is obtained. At temperatures of 60–70°C. a complicated mixture of replacement products results (57, 58).

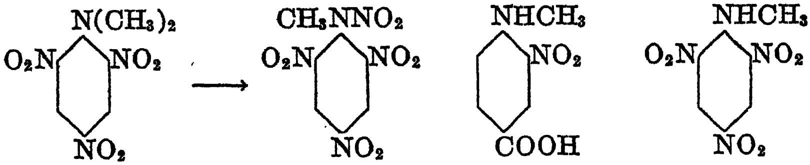
TABLE 3

Nitration of phenolic acids

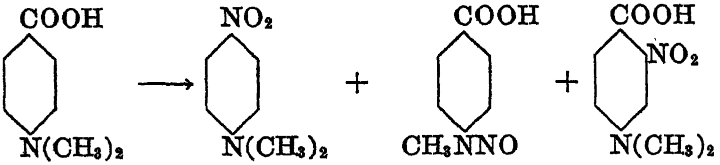
COMPOUND	REAGENT	PRODUCTS	REFERENCES
Salicylic acid	HNO_3 (fuming)	Picric acid	(18)
Anisic acid	HNO_3 (fuming)	Dinitroanisole; 3,5-dinitroanisic acid	(18)
	HNO_3, H_2SO_4	2,4-Dinitroanisole	(66)
	HNO_3, H_2SO_4	3,5-Dinitroanisic acid; 2,4,6-trinitroanisole	(23)
o-Nitrosalicylic acid	HNO_3 (fuming)	*o*-Dinitrophenol; dinitrosalicylic acid	(37)
3,5-Dibromosalicylic acid	HNO_3	6-Bromo-2,4-dinitrophenol	(62)
Dichlorosalicylic acid	HNO_3, CH_3COOH	Dichloronitrophenol	(69)
β-Resorcylic acid	HNO_3	Styphnic acid	(34)
Veratric acid	Hot HNO_3 (1.2)	4-Nitroveratrol	(46)
	HNO_3 (fuming)	3,4,6-Trinitroveratrol	
m-Hemipinic acid		Dinitroveratrol; trinitroveratrol	(33)
Hemipinic acid		Nitrodimethoxybenzoic acid	(88)



Prolonged action of large amounts of nitric acid results in a number of other substances (57):

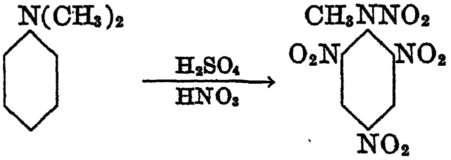


The ease with which the carboxyl group and a methyl group are displaced in *p*-dimethylaminobenzoic acid is illustrated by its reaction with sodium nitrite and hydrochloric acid. The principal products, formed in nearly equal amounts, are as follows (11):

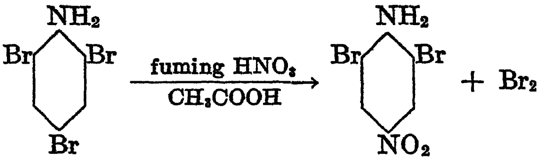


p-Diethylaminobenzoic acid yields corresponding compounds.

One methyl group in dimethylaniline is replaced by a nitro group when the amine is first dissolved in cold sulfuric acid and then nitrated with nitric acid at 40–55°C. (85).



When 2,4,6-tribromoaniline is nitrated with fuming nitric acid in glacial acetic acid, the bromine in the 4-position is replaced by a nitro group (50):



Similar results were obtained with 2-chloro-4,6-dibromo- and 2,6-dichloro-4-bromoanilines, but a chlorine in the 4-position is not replaced.

X. SULFONIC ACIDS

The nitration of sulfonic acids may lead either to normal nitration products or to the replacement by a nitro group of either the sulfonic group, a halogen, or an alkyl group. As might be expected, the products formed depend on the structure of the compound and the experimental conditions. The sulfonic group is readily replaced in the phenolsulfonic acids.

Since numerous examples of these reactions and an extensive bibliography are available in Suter's recent book on the organic sulfur compounds (79), it is not necessary to review them here.

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THE CHEMISTRY OF SODIUM PHOSPHATES

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I. INTRODUCTION

Terrey's statement (125) regarding the chaotic state of our knowledge of metaphosphate chemistry applied, until quite recently, to a wider range of phosphate compositions,—namely, to the pyro- to meta-phosphate range. The term "condensed phosphates" is convenient for this group, and will include both chemical individuals and glasses in this range of Na_2O–P_2O_5 composition. Even the comparatively well-known orthophosphates have recently yielded some little-appreciated points of departure from the simple picture presented by inorganic textbooks. While greater care in thermal analyses and wider appreciation of the utility of the polarizing microscope[1] would have dispelled some of the confusion, a full grasp of phosphate chemistry was hardly to be expected before the application of x-rays to the characterization of the solid phases.

The water-softening power of glasses approximating $NaPO_3$ in P_2O_5 content (i. e., Graham's salt) and of sodium triphosphate has now become well known. However, such products were known chemically for forty to eighty years before any industrial use was made of them, and then it was not water softening but another property which led to the first real industrial application of these compounds.

In the boiler-water field there was need of a *latent* acid to aid in keeping down alkalinity developed by loss of carbon dioxide from carbonates. A substance was required which would be nearly neutral initially, but which would develop acidity in the boiler and thus counteract the alkalinity similarly developed. As is well known, the condensed phosphates react with water, especially at elevated temperatures, to produce acid orthophosphates as the end products. In 1929 Hall and Jackson (48) sought a patent on metaphosphate for this purpose.

The availability of such a phosphate soon led to the discovery that glasses approximating $NaPO_3$ in P_2O_5 content were quite effective in suppressing calcium ions (45).[2] The superiority of the water-soluble glass over the water-soluble crystalline salt was recognized early (46). This Na_2O–P_2O_5 glass was the first water softener whose addition to calcium hard water did not necessarily yield a precipitate, yet could give water practically as soft as distilled water.

The development of such a product led to a re-investigation of the whole range of condensed sodium phosphates. This search was intensified in Europe by the discovery of the calcium-ion-suppressing power of the Trilons (88), i. e., salts of organic polybasic acids such as $(HOOCCH_2)_2NCH_2CH_2N(CH_2COOH)_2$, called Trilon B abroad (or Nullapon B in this country). The Trilons focussed attention on polybasic acids in general, and on polyphosphates in particular, as means of tying up calcium ions in a complex ion. At that time it was still supposed that there might exist a large number of poly-salts in the condensed phosphate region. The patent literature soon contained claims to numerous complex phosphates which were said to be outstanding in water softening, detergent action, wetting, etc. (e. g., 23, 34, 73)[3]. More or less comprehensive studies of the condensed phosphates have since been made in many laboratories.

---

[1] For example, by taking the precaution to characterize the triphosphate hydrate microscopically, as yielding twinned rectangular plates showing extinction 8–9° to the twinning plane, Schwarz (117) proved beyond doubt that he had made $Na_5P_3O_{10}$ in 1895, as has been repeatedly demonstrated recently (12, 60, 95, 101).

[2] As pointed out by Partridge (94), the fact that such glasses could suppress magnesium ion and manganous ion had been known prior to this.

[3] The recent review of Easterwood (31) contains many patent references

This review is an attempt to give a consistent interpretation, primarily from the point of view of the phase rule, of all information now available in the sodium phosphate field. The fundamentals will be stressed more than the practical applications, which have received attention in recent reviews (31, 54, 64, 74, 116).

II. NOMENCLATURE

Metaphosphates differ slightly in classification from other known phosphates, but all condensed phosphates may be regarded as derived from H_3PO_4. They have often been referred to in patent literature as "salts of phosphoric acid poorer in water of constitution than orthophosphoric acid." The basis for this statement is shown in the first two columns of table 1. All except the metaphosphates can also be written in terms of a single general formula, e. g., $Na_{n+2}P_nO_{3n+1}$, as illustrated in the next two columns.

TABLE 1

Nomenclature

DERIVATION OF ACID		NORMAL SALT		FULL NAME OF NORMAL SALT
$nH_3PO_4 - (n-1)H_2O$	$1\ H_3PO_4 - 0\ H_2O$	$Na_{n+2}P_n$- O_{3n+1}	Na_3PO_4	Trisodium orthophosphate
	$2\ H_3PO_4 - 1\ H_2O$		$Na_4P_2O_7$	Tetrasodium pyrophosphate
	$3\ H_3PO_4 - 2\ H_2O$		$Na_5P_3O_{10}$	Pentasodium triphosphate
	$4\ H_3PO_4 - 3\ H_2O$		"$Na_6P_4O_{13}$"	"Hexasodium tetraphosphate"
$nH_3PO_4 - nH_2O$	$3\ H_3PO_4 - 3\ H_2O$	$(NaPO_3)_n$	$Na_3P_3O_9$	Trisodium trimetaphosphate
	$6\ H_3PO_4 - 6\ H_2O$		"$Na_6P_6O_{18}$"*	"Hexasodium hexametaphosphate"

* Some (70, 135) prefer to write the phosphate glasses $Na_2O(NaPO_3)_n = Na_{n+2}P_nO_{3n+1}$. When water of constitution is also present, the formula should be written $(Na,H)_{n+2}P_nO_{3n+1}$. Only when a glass has been made by quenching a melt of $Na_3P_3O_9$ is the formula $(NaPO_3)_n$ acceptable.

Ortho-, pyro-, and tripolyphosphates could quite properly be called mono-, di-, and triphosphates, respectively. Obviously the prefixes "ortho" and "pyro" are too well established to permit changing to the systematic nomenclature. The prefix "tripoly" for the P_3O_{10} ion, though poor chemically, has the advantage of directing one's thoughts more clearly away from Na_3PO_4, which is so often called *tri*sodium phosphate. "Tripoly" is so widely used in industry that it seems likely to prevail, but the shorter prefix will be used here.

It has been common practice to omit the prefix "ortho" when discussing orthophosphates. This is no longer acceptable, for it will inevitably lead to confusion.

The quotation marks in table 1 indicate that "tetraphosphate" and "hexametaphosphate" are not simple chemical individuals. Both will be discussed more fully later.

Another confusing expression needs clarification. One often speaks of glassy metaphosphate or metaphosphate glass. Neither term is free from objections but the former is especially bad, for it implies, particularly when set alongside the expression "crystalline metaphosphate," that a glass of the $NaPO_3$ composition is a particular compound. It is not a compound, but merely a supercooled liquid; not only are its molecular units constrained in some random arrangement, but it probably contains a range of molecular sizes and structures. In the interest of clarity such terms should be avoided.

III. PHASE DIAGRAMS

Perhaps the most satisfactory way to get a comprehensive picture of phosphate chemistry and physics is by means of phase diagrams. While the data available are rather limited, nevertheless they suffice to answer many important questions n this complex field.

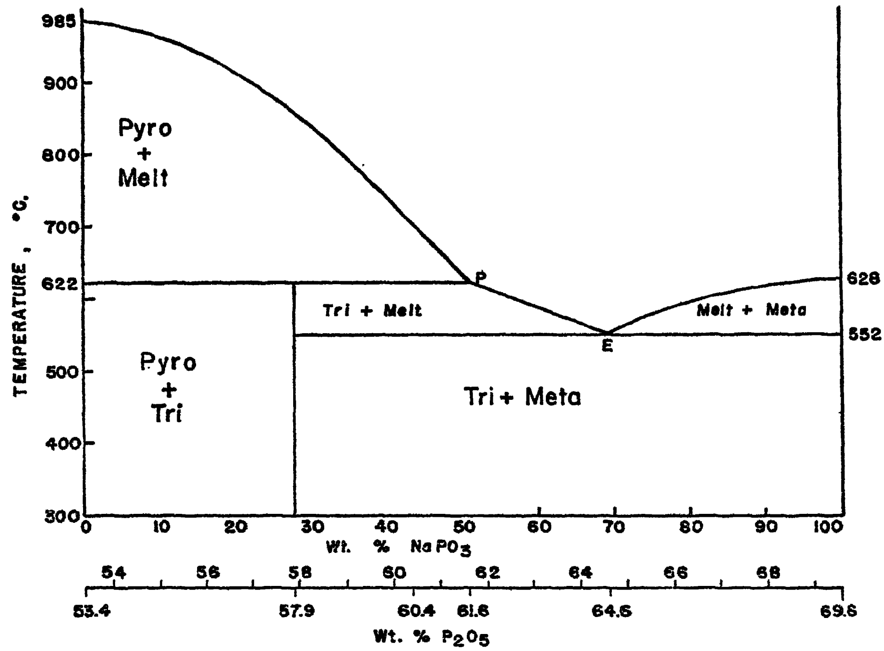

FIG. 1. The $Na_4P_2O_7$–$NaPO_3$ system

A. BINARY SYSTEM Na_2O–P_2O_5

Data are available only for the $Na_4P_2O_7$–$NaPO_3$ portion of the Na_2O–P_2O_5 system, and these are far from complete. Data based on heating curves plus microscopic and x-ray observations are shown in figure 1, which is based on the work of Morey and Ingerson (83) and of Partridge and coworkers (95).[4] The ordinate is temperature in degrees Centigrade and the abscissa is composition expressed in terms of per cent $NaPO_3$, a procedure which is somewhat misleading. A second (per cent P_2O_5) abscissa scale has therefore been added. This diagram does not extend below 300°C., because very little if any change occurs at lower temperatures. Also not indicated are some solid–solid transformations in the range 400–550°C. for the crystalline salts $Na_4P_2O_7$, $Na_5P_3O_{10}$, and $NaPO_3$.

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[4] It is recognized that the vertical lines at compositions $Na_4P_2O_7$, $Na_5P_3O_{10}$, and $NaPO_3$ cannot represent the true state of affairs, but the necessary solid–solid intersolubilities are lacking for showing them correctly as one-phase areas of finite width.

The meaning of figure 1 may be illustrated in a few cases. Whereas the crystalline salts $Na_4P_2O_7$ and $NaPO_3$ melt to a liquid of the same Na_2O/P_2O_5 ratio, $Na_5P_3O_{10}$ *decomposes as it melts* at 622°C. to give crystals close to $Na_4P_2O_7$ in composition and a liquid indicated by point *P* (61.6 per cent P_2O_5); continued heating simply decreases the amount of $Na_4P_2O_7$-like solid phase until the last crystal disappears at 860°C., giving a melt containing 57.9 per cent P_2O_5. In theory these changes are simply reversed if such a melt is cooled, but in practise the *manner* of cooling has an important bearing on the chemical composition of the product. The reason for this is that reactions involving solid phases are slow and that the liquids involved in phosphate systems are very viscous so that equilibrium is difficult to realize. Sufficiently rapid chilling will give 100 per cent glass. Moderately rapid chilling of a melt made from $Na_5P_3O_{10}$ gives no crystals of $Na_5P_3O_{10}$, but rather fine crystals of $Na_4P_2O_7$ imbedded in a sodium phosphate glass. Slower cooling gives a mixture of $Na_4P_2O_7$, $Na_5P_3O_{10}$, $NaPO_3$, and more or less sodium phosphate glass.

How then is pure sodium triphosphate to be made? Figure 1 indicates that $Na_5P_3O_{10}$ can form in the Na_2O–P_2O_5 system only under conditions that demand reaction between solids. Since such reactions are slow, the reactants must obviously be very finely divided and intimately mixed. This condition can be met either by grinding the reactants together cold, or by quick chilling of the melt. In either case $Na_5P_3O_{10}$ of high purity can be made by heating the intimate mixture for a considerable time at temperatures near 500°C.

Melts containing more than 50 per cent "$NaPO_3$" (61.6 per cent P_2O_5) do not crystallize at all if a fast cooling rate is used, the product being a glass, as recognized by Parravano and Calcagni (92) in 1908. If the cooling rate is about 5°C. per minute, these mixtures supercool about 100°C. (95) before beginning to crystallize; products so cooled to room temperature contain substantial amounts of glass. No one has succeeded in getting equilibrium data by cooling melts in this portion of the Na_2O–P_2O_5 diagram. Because of this fact, investigators have found it easier to determine equilibria by observing the thermal changes upon heating intimate mixtures of various Na_2O–P_2O_5 compositions.

In the "pyro-rich" part of the diagram (61.5–53.4 per cent P_2O_5) there is less tendency toward supercooling the nearer the composition approaches that of $Na_4P_2O_7$. Hence fairly reliable freezing-point data may be obtained (e.g., they are only 10–30°C. low for a cooling rate of 5°C. per minute). Of particular interest is the composition 60.4 per cent P_2O_5, corresponding to the formula "$Na_6P_4O_{13}$" given for the so-called sodium tetraphosphate. While this point has been marked on the P_2O_5 scale for the reader's benefit, the phase diagram gives no evidence for the existence of such a compound. Perhaps it would be well to consider the implications of figure 1 for a melt of the "tetraphosphate" P_2O_5 content, and to simplify the description it may be assumed that equilibrium is attained.

Cooling such a melt from, say, 800°C. first leads to crystals of $Na_4P_2O_7$, which begin forming at 705°C. As more $Na_4P_2O_7$ crystallizes, the liquid composition moves toward point P, which is reached at a temperature of about 622°C. On passing below this peritectic temperature, all $Na_4P_2O_7$, by interaction with the melt, is transformed to $Na_5P_3O_{10}$.[5] On further cooling, more $Na_5P_3O_{10}$ crystallizes and the liquid composition moves toward the eutectic E. When the temperature falls below 550°C., the remaining eutectic liquid solidifies. The final mixture is composed of 78.3 per cent triphosphate and 21.7 per cent metaphosphate distributed as follows:

60.6 per cent $Na_5P_3O_{10}$ crystallized above 550°C.
17.7 per cent $Na_5P_3O_{10}$ crystallized below 550°C.
21.7 per cent $NaPO_3$ crystallized below 550°C.

Reversal of thermal changes simply reverses these changes. Thus, products marketed as sodium "tetraphosphate" are mixtures. This has been demonstrated in other ways (10, 22).

Any failure to attain equilibrium conditions in manufacture will complicate the mixture. It may easily cool too fast to convert all $Na_4P_2O_7$ to $Na_5P_3O_{10}$ and this would necessitate more P_2O_5-rich material (either $NaPO_3$ or Na_2O–P_2O_5 glass or both). If the product is made from reactants such as Na_2HPO_4 and NaH_2PO_4, incomplete dehydration may leave acid salts, such as $Na_3HP_2O_7$, $Na_4HP_3O_{10}$, or $(Na,H)_{n+2}P_nO_{3n+1}$.

Before leaving the "tetraphosphate" composition it should be recalled that reaction products in these systems are usually examined cold. Conceivably $Na_6P_4O_{13}$ and/or other higher polyphosphates may be found to have narrow thermal ranges of stability at elevated temperatures, by analogy with sodium pyrosilicate, $Na_6Si_2O_7$ (137), which is stable above 402°C. but decomposes at ordinary temperatures.

B. BINARY SYSTEM H_2O–P_2O_5

This is even less understood than the Na_2O–P_2O_5 system. The phosphoric acids are very viscous, so that equilibria are attained *very* slowly, the usual result of cooling being a glass. Furthermore, they are hygroscopic and attack ordinary glass vessels. All that can be stated relative to the *equilibrium* H_2O–P_2O_5 diagram is that the melting points of three compounds are known, as follows:

COMPOUND	MELTING POINT	REFERENCE
	°C.	
H_3PO_4	42.4	(107, 122, 132)
$H_4P_2O_7$	*ca.* 61	(41)
P_2O_5 (tetragonal?)	580	(57)

By adding P_2O_5 in small increments to water until the mixture contained 84 per cent P_2O_5, Rakuzin and Arsenev (102) obtained a crystalline product after 5 days. It gave a melting point of 34°C. While this might have been a new compound such as $H_5P_3O_{10}$, no optical or x-ray data were given to show that it differed from H_3PO_4 or $H_4P_2O_7$.

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[5] Practically speaking, this transformation is very difficult to realize quantitatively.

Satisfactory equilibrium data in the H_2O–P_2O_5 system, especially between H_3PO_4 (72.4 per cent P_2O_5) and P_2O_5, will probably not be obtainable by cooling methods. Perhaps one could induce crystallization by the supercooling procedure applied by Schierloh (113) to glycerol. The sample is chilled at some very low temperature for a while and then held at a higher temperature, but below the complete melting point, for crystallization and stabilization. The melting data found upon warming such a stabilized system will probably be close to equilibrium values.

The behavior of P_2O_5 shed some light on the polymer-forming tendency of the complex phosphates. According to Hill, Faust, and Hendricks (57) there are three polymorphic forms. The sublimed P_2O_5 can be resublimed completely at 250°C. and 15 mm. (42) and has a hexagonal crystal structure with four P_2O_5 units in the unit cell. It melts in a closed system at 422°C. but quickly resolidifies in the metastable orthorhombic form; this form melts at about 558°C. and slowly resolidifies to the stable or tetragonal (?) form. The vapor pressure of hexagonal P_2O_5 is highest, that of orthorhombic is much lower, and that of tetragonal is still lower. Hexagonal P_2O_5 reacts violently with water, as everyone knows; orthorhombic P_2O_5 reacts much more slowly (crystals may persist after an hour's contact with water); tetragonal P_2O_5 falls *between* the two. While the x-ray data (57) have not been completely analyzed, the stable (or tetragonal ?) form of P_2O_5 is suspected of being a three-dimensional polymer; it can be superheated to 700°C. for a short time without melting, probably because it cannot depolymerize rapidly enough to melt immediately.

C. TERNARY SYSTEM $NaOH$–H_3PO_4–H_2O

Here again only fragmentary data are available. At room temperature and a pressure of 1 atm., the composition triangle is fairly complete for mixtures of NaOH, H_3PO_4, and H_2O as indicated in figure 2. This graph is based on the data of d'Ans and Schreiner (25) at 25°C. and those of Menzel and von Sahr (82) at 20°C.; the latter have been crudely translated to 25°C. The data are plotted on a right triangle, per cent P_2O_5 along the ordinate and per cent Na_2O along the abscissa; per cent H_2O is found by difference.

A few reference lines are dotted across the triangle to aid in showing chemical relationships. For example, one might ask why not plot the data in terms of NaOH, H_3PO_4, and H_2O, the components mixed. This would reduce the triangle to the position shown by the dotted line running from the NaOH to the H_3PO_4 composition. Now we see that one solid obtained (point *G*, $Na_3PO_4 \cdot \frac{1}{2}H_2O$) lies outside this triangle, so that the results could not be properly expressed in terms of $NaOH$–H_3PO_4–H_2O unless negative values were used. This needless complication is avoided if the data are plotted in terms of Na_2O, P_2O_5, and H_2O. Certain key (molar) ratios of $Na_2O:P_2O_5$ are marked by dotted lines from 100 per cent H_2O (i. e., 0 per cent Na_2O, 0 per cent P_2O_5) to the Na_2O–P_2O_5

side of the triangle. Definitely delineated are one-, two-, and three-phase areas as follows:

1. One-phase areas
    a. Homogeneous solution (liquid)
    b. Seven solid phases (indicated by squares)
2. Six two-phase areas (marked by tie lines)
3. Six three-phase areas (two solids in equilibrium with one liquid solution)

1. Salts not decomposed by water

If we put $NaH_2PO_4 \cdot 2H_2O$ or $Na_2HPO_4 \cdot 12H_2O$ into water, heat the system to produce complete solution, and cool back to 25°C., we shall (at the proper concentrations) recover some of the crystals in a form identical with that put into solution, the remainder being in the saturated solution. This operation can be repeated, using the product of the first crystallization, with the same result. This is possible because the Na_2O/P_2O_5 ratio is the same in the mother liquor as in the solid. While there is nothing unusual about such behavior, many salts, including some sodium phosphates, show a different behavior.

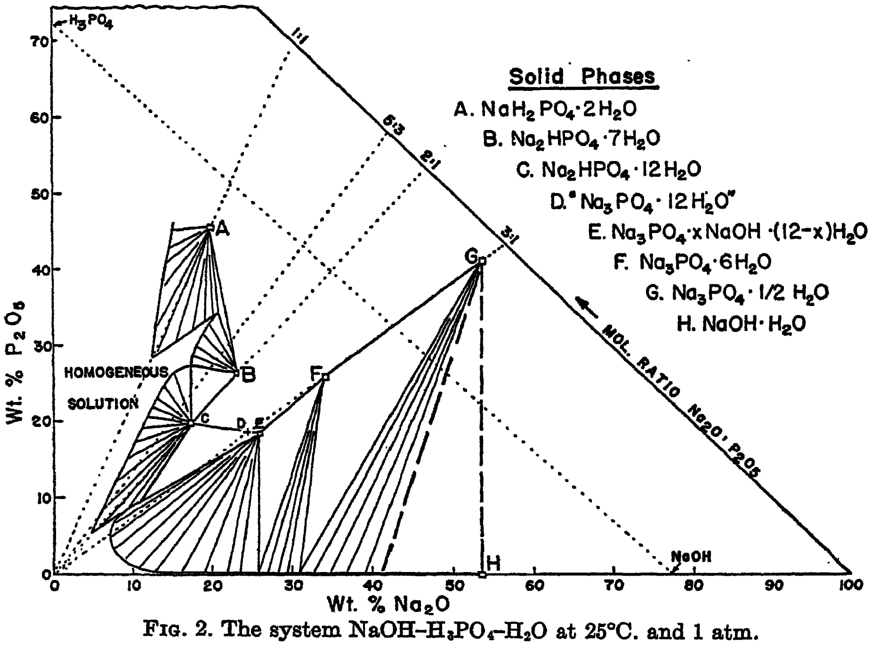

FIG. 2. The system $NaOH–H_3PO_4–H_2O$ at 25°C. and 1 atm.

We could express the above facts in another way. Since the Na_2O/P_2O_5 ratio is the same in crystal as in saturated solution for all NaH_2PO_4 hydrates, NaH_2PO_4 in water (or Na_2HPO_4 in water) can be treated as a two-component system. Examination of Seidell's handbook (118) reveals the following hydrates as the saturating solids:

NaH_2PO_4		Na_2HPO_4	
Hydrate	Temperature range	Hydrate	Temperature range
	°C.		°C.
$2H_2O$	0–41	β $12H_2O$	0–30
$1H_2O$	41–58	α $12H_2O$	30–35
$0H_2O$	Above 58	$7H_2O$	35–48
		$2H_2O$	48–95
		$0H_2O$	Above 95

A single solubility curve can be drawn for either salt as a function of temperature from 0° to 100°C. or above.

Two crystalline phases having the composition $Na_2HPO_4 \cdot 12H_2O$, which are stable in contact with saturated solution, were detected by solubility measurements (49); the α-form is stable at 30–35°C. and the β-form at 0–30°C.

2. Salts decomposed by water

The salt $Na_2HPO_4 \cdot 7H_2O$ can be made not only by crystallization from water at 35–48°C., as indicated above, but can also be crystallized at 25°C. from concentrated solutions containing a suitable excess of H_3PO_4 (see figure 2). Even in such acid solutions the addition of sufficient water will decompose the heptahydrate and crystallize the dodecahydrate. Furthermore, crystals of heptahydrate cannot be purified by recrystallization from water at 25°C., for only the dodecahydrate will be obtained.

A similar but reversed situation exists with respect to the salts $Na_3PO_4 \cdot \frac{1}{2}H_2O$ and $Na_3PO_4 \cdot 6H_2O$. These require an excess of alkali to crystallize at 25°C. A line drawn from a point representing the total composition of a system where $Na_3PO_4 \cdot \frac{1}{2}H_2O$ is in equilibrium with its saturated solution to 100 per cent water (point 0,0 in figure 2) shows how a small amount of water will yield $Na_3PO_4 \cdot 6H_2O$ + saturated solution, while a larger amount will lead to the alkaline complex E. Thus it is impossible to recrystallize the two lower hydrates of Na_3PO_4 from water at 25°C. Either one would yield the alkaline complex.

For most of the crystalline phases in the Na_2O–P_2O_5–H_2O system, the Na_2O/P_2O_5 ratio in the crystal remains constant in spite of (limited) variations of this ratio in the total system and in the saturated solution. This is not true of the crystalline phase represented by square E in figure 2; though plotted as a point this refers to a series of solid solutions which are always richer in Na_2O than $Na_3PO_4 \cdot 12H_2O$, the formula often written for the compound obtained by crystallizing Na_3PO_4 from water at ordinary temperatures. By a study of the isobaric absorption of water vapor at a partial pressure of 10 mm. of mercury, Menzel and von Sahr (82) obtained a crystal analyzing as $Na_3PO_4 \cdot 12H_2O$ and giving an x-ray pattern very similar to that of the more complex hydrate containing the extra sodium hydroxide.[6] The instant one puts this true dodecahydrate into a limited quantity of liquid water, it is converted either to a mixture of $Na_2HPO_4 \cdot 12H_2O$ and $Na_3PO_4 \cdot xNaOH \cdot (12 - x)H_2O$ in contact with solution or to the latter salt alone in contact with solution. The isobaric decomposition curves of these two "dodecahydrates" (82) show that the extra sodium hydroxide stabilizes the product; thus at a partial pressure of 10 mm. of water, dehydration to the octahydrate occurs at 26–26.5°C. for the more alkaline salt and at 15–15.5°C. for the other.

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[6] Either Na_3PO_4, $Na_3PO_4 \cdot \frac{1}{2}H_2O$, or $Na_3PO_4 \cdot 6H_2O$ might be used for this vapor hydration, according to Menzel and von Sahr. Addition of water vapor to the hexahydrate leads first to the octahydrate, and then finally at sufficiently low temperature (less than 15°C.) to the

The solid of 3.22 Na_2O/P_2O_5 ratio is the only one that can exist in equilibrium with a saturated solution having the same Na_2O/P_2O_5 ratio. Solids of lower ratio give more acid solutions; those of higher ratio give more alkaline solutions. Thus a tertiary sodium orthophosphate purified by several recrystallizations from water will have an Na_2O/P_2O_5 ratio of 3.22/1. The lowest value of x observed in products crystallized from water was 3.11:1; the original mixture had a 2.80:1 ratio and the mother liquor a 2.69:1 ratio. The highest value observed was 3.25 (large excess of sodium hydroxide in solution). Thus x varies from 0.1 to 0.25.

Analyses of the commercial dodecahydrate have shown repeatedly (e. g., 67, 82, 101) that there is extra alkali in so-called "$Na_3PO_4 \cdot 12H_2O$", the Na_2O/P_2O_5 ratio found varying from 3.14 to 3.20. Menzel and von Sahr (82) found one product to have the ratio 2.98:1, but microscopic examination of the crystals revealed two solid phases: namely, $Na_2HPO_4 \cdot 12H_2O$ and $Na_3PO_4 \cdot xNaOH \cdot (12 - x)H_2O$.

The formula of this tertiary dodecahydrate is written with the x to signify variability. The x-ray data have not been completely interpreted, but they suggest that the extra sodium hydroxide goes into the crystal lattice in the place of some of the water of crystallization, hence the $(12 - x)H_2O$.

It is natural to compare this series of solid solutions for tertiary sodium orthophosphate with those observed among the corresponding calcium phosphates, particularly the apatites. For instance, we may tabulate formulas as follows:

$(Ca_3P_2O_8)_3 \cdot CaF_2$ or $Ca_5P_3O_{12}F$	Apatite
$(Ca_3P_2O_8)_3 \cdot CaCl_2$ or $Ca_5P_3O_{12}Cl$	Chlorapatite
$(Ca_3P_2O_8)_3 \cdot Ca(OH)_2$ or $Ca_5P_3O_{12}(OH)$	Hydroxyapatite
$(Ca_3P_2O_8)_3 \cdot 2H_2O$?	Aquo-apatite?
$Ca_3P_2O_8 \cdot Ca(OH)_2 \cdot 4H_2O$	Isoclasite
$(Na_3PO_4)_2 \cdot NaF \cdot 19H_2O$	Wrongly called "$Na_3PO_4 \cdot 10H_2O$"
$(Na_3PO_4)_4 \cdot NaOH \cdot 47H_2O$ to $(Na_3PO_4)_{10} \cdot NaOH \cdot 119H_2O$	Series of solid solutions

Writing the formulas of the apatites as above is misleading for they suggest definite chemical individuals, whereas actually the compositions indicated are

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dodecahydrate. All of these salts were characterized by analysis for water and by having a distinctive x-ray pattern.

to be taken as illustrations of extremes in systems of solid solutions. For example, the fluoride of apatite can be partially or wholly replaced by one or more radicals such as chloride, carbonate, hydroxy, etc. without destroying or changing greatly the apatite lattice (32). It has even been suggested that there exists an aquo-apatite and a series of solid solutions between it and hydroxyapatite (56), but this is doubted by others (32). Among known sodium compounds is a fluoro derivative (78), but it is hydrated; it was once erroneously reported as $Na_3PO_4 \cdot 10H_2O$ because analyses were not made for fluorine and the water content was close to that for the decahydrate. The extremes of the series of solid solutions denoted by E in figure 2 can be written as $(Na_3PO_4)_4 \cdot NaOH \cdot 47H_2O$ and $(Na_3PO_4)_{10} \cdot NaOH \cdot 119H_2O$. All have the same microscopic appearance and essentially the same x-ray pattern (82), just as do certain series of solid solutions that have been studied among the apatites.

From the foregoing it is evident that we cannot properly speak of the solubility of Na_3PO_4 in water. This is certainly true in the temperature range where the variable alkaline hydrate is the stable saturating phase. In short, Na_3PO_4–H_2O cannot be treated as a two-component system. The fact that the saturated solution and saturating solid may differ from each other and from the initial crystal in Na_2O/P_2O_5 ratio probably accounts for the discordance of Na_3PO_4 "solubility" data. The data quoted by Seidell (118) plus those of Menzel and von Sahr (82) suggest that the saturating phases are as follows:

$Na_3PO_4 \cdot xNaOH \cdot (12 - x)H_2O$	−0.5° to about 43°C.
$Na_3PO_4 \cdot 8H_2O$ (or $7H_2O$?)	43° to 70°C.
$Na_3PO_4 \cdot 6H_2O$	70° to 121°C.
$Na_3PO_4 \cdot \frac{1}{2}H_2O$ (or $1H_2O$?)	121° to 212°C.
Na_3PO_4	Above 212°C.

Four stages of hydration are clearly evident from breaks in the "solubility" curve, and a fifth seems to be present near 43°C. How many hydrates exist in the 6–8 H_2O range is not yet certain and there is still argument over the lower hydrate ($\frac{1}{2}$ or 1 H_2O?). Whether the lower hydrates (8, 6, and $\frac{1}{2}$ H_2O) are variable remains to be demonstrated by careful systematic studies, but Menzel and von Sahr (82) imply that the hexahydrate and the hemihydrate are simple chemical individuals.

3. Discussion

In the $NaOH$–H_3PO_4–H_2O system (figure 2) no evidence was found for pyro-, tri-, or metaphosphates in spite of the fact that the region involving these Na_2O/P_2O_5 ratios was at least partially explored. This is contrary to the behavior of the silicates or vanadates. In the Na_2O–V_2O_5–H_2O system (66, 77, 105), for example, addition of vanadic acid to Na_3VO_4 to reduce the Na_2O/V_2O_5 ratio to about 2 gives crystals of hydrated $Na_4V_2O_7$, not Na_2HVO_4; further addition of V_2O_5 to a Na_2O/V_2O_5 ratio near 1 gives crystals of hydrated $NaVO_3$, not NaH_2VO_4. The different behavior of the phosphates suggests that the orthophosphates are relatively more stable than the corresponding vanadates.

As Hendricks (55) has pointed out, a further indication of orthophosphate stability is to be found in the phosphate minerals. The one hundred and sixty phosphate and arsenate minerals listed in Dana's *Textbook of Mineralogy* (24) are almost exclusively ortho derivatives.

Still further evidence is given by dilute solutions of phosphates in which pyro-, tri-, and metaphosphates tend to hydrate to ortho. In particular, Bamann (5) showed that in the presence of a suitable catalyst at pH 8–9 even pyrophosphate hydrates readily to orthophosphate at 37°C.

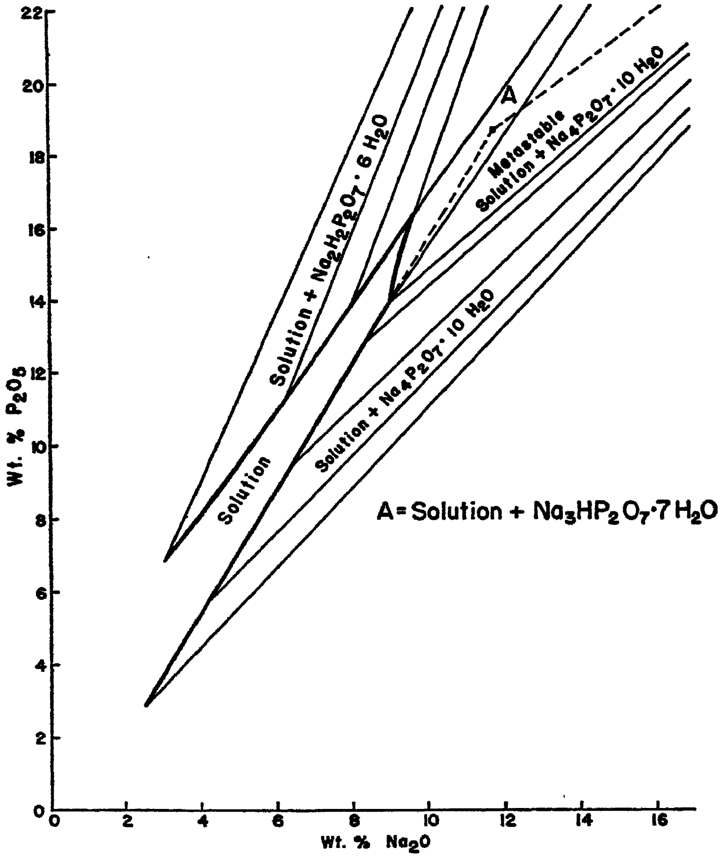

FIG. 3. The system $NaOH\text{-}H_4P_2O_7\text{-}H_2O$ at 20°C. and 1 atm.

It is probably safe to assume that pyrophosphates have either a very limited or even no stable existence in the $Na_2O\text{-}P_2O_5\text{-}H_2O$ system; tri- and metaphosphates are even less stable[7] than pyrophosphates.

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[7] Further discussion of the stability of condensed sodium phosphates will be found in the last section, page 176.

D. TERNARY SYSTEM $NaOH-H_4P_2O_7-H_2O$

While this system is certainly metastable with respect to orthophosphates over part of the range and may be metastable over the whole range, yet pyrophosphates change to ortho slowly enough in the absence of catalysts at ordinary temperatures to permit evaluation of some of the metastable equilibria. Selva's data for 20°C. and 1 atm. (119) are shown in figure 3 for the region 1:1 to 2:1 Na_2O/P_2O_5 ratios; the data are plotted on a right triangle; per cent H_2O is found by difference.

The homogeneous-solution region is much smaller than for orthophosphates. The solid phases which crystallize from solution are $Na_4P_2O_7 \cdot 10H_2O$, $Na_3HP_2O_7 \cdot 7H_2O$, and $Na_2H_2P_2O_7 \cdot 6H_2O$. Again, nothing appears at the 5:3 ratio, which we have in $Na_5P_3O_{10}$, and the salt at the 1:1 ratio is not $NaPO_3$, but $Na_2H_2P_2O_7$. This system could surely be studied in more alkaline systems. The homogeneous-solution branch should bend around and go toward 40.5 per cent Na_2O at 0 per cent P_2O_5, where saturated sodium hydroxide solution would be in equilibrium with solid $NaOH \cdot H_2O$. $Na_4P_2O_7$ or possibly other, lower hydrates of $Na_4P_2O_7$ would probably put in their appearance. Probably more acid systems would be difficult to study, owing to rapid reversion to orthophosphates.

This system has also been studied to the same extent at 30°C. and 1 atm. Here the system differs only in having a slightly larger homogeneous-solution area and in having $Na_2H_2P_2O_7$ and $Na_3HP_2O_7 \cdot H_2O$ as the acid pyrophosphate saturating solids, instead of $Na_2H_2P_2O_7 \cdot 6H_2O$ and $Na_3HP_2O_7 \cdot 7H_2O$, respectively.

Textbooks (e. g., 124) often imply that the salt $Na_3HP_2O_7$ or its hydrates do not exist, and the older literature is (79) certainly confusing on this point. Selva's data suggest reasons for the difficulty. In the first place the range of existence of $Na_3HP_2O_7$ is quite narrow at 20–30°C. Then, a metastable $Na_4P_2O_7 \cdot 10H_2O$–solution equilibrium can be realized without any $Na_3HP_2O_7 \cdot 7H_2O$ being present (figure 3). Also, when drawn on a very large scale, Selva's data suggest that $Na_3HP_2O_7 \cdot 7H_2O$ alone may be unable to exist in equilibrium with a saturated solution of the same Na_2O/P_2O_5 ratio, but require slightly more acid solutions. Contrast with this situation the fact that both $Na_4P_2O_7$ and $Na_2H_2P_2O_7$ can be recrystallized directly from water[8] giving the following salts:

$Na_2H_2P_2O_7 \cdot 6H_2O$ (119)	−0.7°C. to 27°C.
$Na_2H_2P_2O_7$ (119)	Above 27°C.
$Na_4P_2O_7 \cdot 10H_2O$ (80, 81, 119)	−0.4° to 79°C.
$Na_4P_2O_7$ (81)	Above 79°C.

It is not surprising that $Na_4P_2O_7$ and $Na_2H_2P_2O_7$ are better known than $Na_3HP_2O_7$.

Roborgh (106) found $Na_4P_2O_7$ more soluble in glycerol (specific gravity 1.2303, 98 per cent) than in water, as indicated below:

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[8] This ignores the tendency of $Na_2H_2P_2O_7$ to hydrolyze or revert to NaH_2PO_4, which is the true equilibrium salt. Similarly, $Na_4P_2O_7$ probably tends to revert to Na_2HPO_4.

SOLVENT	SOLUBILITY AT 20°C.	SOLID PHASE
	per cent	
H_2O	5.22	$Na_4P_2O_7 \cdot 10H_2O$ (81, 119)
Glycerol	8.8	

E. TERNARY SYSTEMS INVOLVING TRI- OR METAPHOSPHATES

Although the study would probably be more difficult, owing to more rapid reversion toward orthophosphates, a similar study on sodium triphosphates, and perhaps even on sodium metaphosphates, could be made. However, very little such information is available.

A sample of commercial sodium triphosphate, twice recrystallized from water in this laboratory, gave a solubility of 14.5 per cent at about 26°C. (101). Ingerson and Morey (60) found the solubility to be 16.5 per cent at about 25°C., the solid phase being $Na_5P_3O_{10} \cdot 6H_2O$.

The common, water-soluble, crystalline sodium metaphosphate, believed by some to be trimeta $Na_3P_3O_9$, is presumably somewhat more soluble, but quantitative data are lacking. Wiesler (134) gave the solubility as 18 per cent in "cold water", without specifying the temperature. The saturating solid is presumably triclinic $Na_3P_3O_9 \cdot 6H_2O$ (or $NaPO_3 \cdot 2H_2O$).

Na_2O–P_2O_5 glasses of P_2O_5 content near that of $NaPO_3$ dissolve to a much greater extent. Strictly speaking, one cannot measure the solubility of a glass, for a solubility determination requires the presence of a definite saturating solid phase in equilibrium with the solution. However, it is useful to know that a clear syrupy solution containing 50–70 per cent of such a glass can be made (20, 64, 121).

F. CRYSTALLINE SODIUM PHOSPHATES

The number of crystalline species in the Na_2O–P_2O_5–H_2O system is evidently quite large. Table 2 lists twenty-nine salts which may be encountered at ordinary temperatures. At temperatures above 300°C. four other forms of $Na_4P_2O_7$ are known (95). It is to be expected that a few more will be found when further systematic studies are made, e. g., $Na_4HP_3O_{10}$ in other levels of hydration than the one listed in table 2. Since any heat treatment may convert the more condensed to less condensed phosphates, this abundance of possible compounds complicates the characterization of such products.

IV. HARDNESS SUPPRESSION

One reason for interest in condensed phosphates is that they possess ability to lock up calcium and magnesium (also ferric) ions and reduce their chance of precipitating soap or other fatty detergents. Even though we limit our attention to one kind of hardness (e. g., calcium) and to one particular soap (e. g., NaMy = sodium myristate), we have increased the complexity of the system to five components Na_2O–CaO–P_2O_5–HMy–H_2O. Very useful information to have on such a system would be the boundary of the homogeneous-solution region, together with the composition of the solid phase in equilibrium with each saturated solution. Needless to say, such data are not available, but their utility and significance can be demonstrated by means of a simpler, analogous system for which data are available. A striking example among three-component systems is $KI–HgI_2–H_2O$.

TABLE 2

Sodium phosphates

TYPE	NORMAL SALT	ACID SALTS	
Ortho	Na_3PO_4	Na_2HPO_4	NaH_2PO_4
	$Na_3PO_4\cdot 0.5H_2O$	$Na_2HPO_4\cdot 2H_2O$	$NaH_2PO_4\cdot H_2O$
	$Na_3PO_4\cdot 6H_2O$*	$Na_2HPO_4\cdot 7H_2O$	$NaH_2PO_4\cdot 2H_2O$
	$Na_3PO_4\cdot 8H_2O$*	$Na_2HPO_4\cdot 12H_2O$ (α)	
	$Na_3PO_4\cdot x NaOH\cdot (12 - x)H_2O$	$Na_2HPO_4\cdot 12H_2O$ (β)	
Pyro	$Na_4P_2O_7$	$Na_3HP_2O_7$	$Na_2H_2P_2O_7$
	$Na_4P_2O_7\cdot 10H_2O$	$Na_3HP_2O_7\cdot H_2O$	$Na_2H_2P_2O_7\cdot 6H_2O$
		$Na_3HP_2O_7\cdot 7H_2O$	
Tri	$Na_5P_3O_{10}$ I	$Na_4HP_3O_{10}\cdot 1.5H_2O$†	
	$Na_5P_3O_{10}$ II		
	$Na_5P_3O_{10}\cdot 6H_2O$		
	$Na_5P_3O_{10}\cdot 8H_2O$‡		
Meta§	$NaPO_3$ I		
	$NaPO_3$ II		
	$NaPO_3$ III		
	$NaPO_3\cdot 2H_2O$		

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* Hall (44) and Ingerson and Morey (60) speak of $Na_3PO_4\cdot 7H_2O$.

† Obtained by Huber (58) by adding alcohol to an aqueous acetic acid solution of $Na_5P_3O_{10}$.

‡ Both Huber (59) and Bonneman-Bémia (13) claim that this is formed by direct crystallization from water; it readily loses water in air to become the hexahydrate.

§ $NaPO_3$ I is the soluble crystalline variety, Boullé's A (15, 16) and Nylén's $Na_3P_3O_9$. $NaPO_3$ II is an insoluble crystalline variety, Boulé's B. $NaPO_3$ III is also insoluble and crystalline but does not correspond to any form reported by Boullé. $NaPO_3\cdot 2H_2O$ is the hydrated form of $NaPO_3$ I.

A. SYSTEM $KI–HgI_2–H_2O$

This system, illustrated for 20°C. in figure 4 (29, 90, 100), has a very large homogeneous-solution area, all of it requiring excess potassium iodide over the 1:1 molar ratio required to form the complex salt $KHgI_3\cdot H_2O$. Thus mercuric iodide, whose water solubility is comparable to that of calcium carbonate, is very soluble in concentrated potassium iodide solutions, owing to the formation of a soluble complex. Formation of the complex may be expressed by means of equations as follows:

$$\begin{array}{rl} HgI_2(\text{solid}) \rightleftharpoons & HgI_2\ (\text{dissolved}) \rightleftharpoons Hg^{++} + 2I^- \\ & + \\ KI \rightleftharpoons K^+ + & I^- \\ & \updownarrow \\ & HgI_3^- \end{array}$$

This reduces the Hg^{++} concentration below that required to exceed the solubility product and hence no mercuric iodide precipitates.

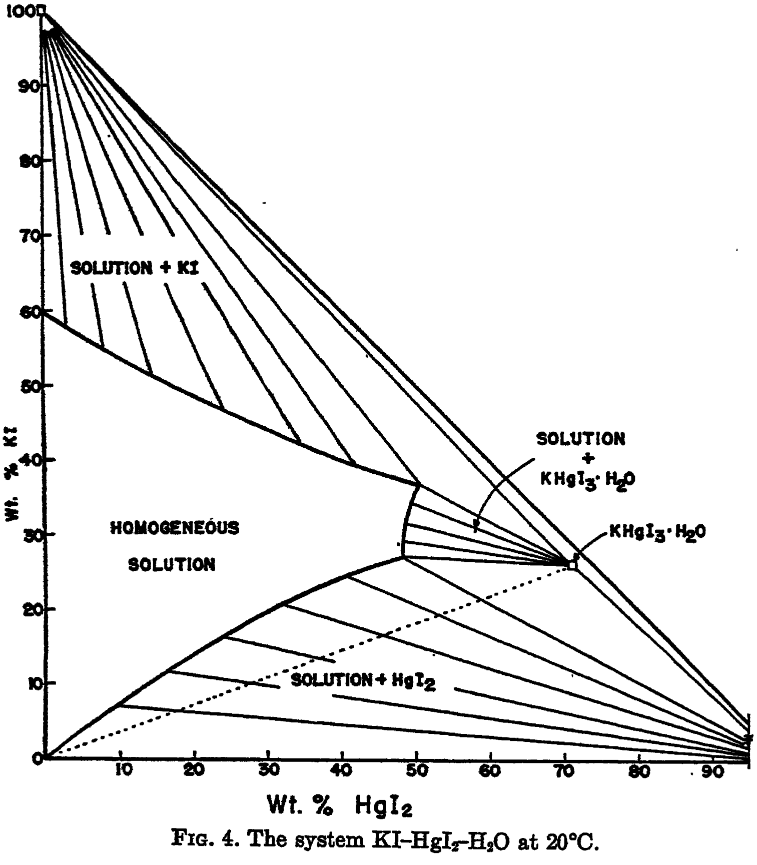

Fig. 4. The system $KI-HgI_2-H_2O$ at 20°C.

B. SUPPRESSION OF Ca^{++}, Mg^{++}, AND Fe^{+++}

While the homogenous-solution region has not been determined in the more complex, practical systems, e. g., $Na_2O-CaO-P_2O_5$–fatty acid–H_2O, nevertheless some parts of this boundary have been located under certain particular conditions. However, the data cannot be presented in the form of a phase diagram.

As a measure of the suppression of calcium, magnesium, or ferric ions it has been customary to determine the minimum amount of the sodium phosphate in question which will (*1*) dissolve a precipitate (101, 126) or more often (*2*) prevent precipitation (4, 22, 36, 40, 52, 59, 101, 108, 110, 126) of some calcium.

magnesium, or iron salt. In spite of the abundance of well-known Fe^{+++} complexes (e. g., $Fe(CN)_6^{---}$), the phosphates seem less effective against ferric iron than against calcium or magnesium (121, 126). It has become evident, however, that complexes are more readily formed with magnesium than with calcium. Thus, precipitation of magnesium with calcium as oxalate can be prevented by the addition of a sufficient excess of ammonium oxalate. This trend is confirmed by the data of Hall (47), who found that the addition of 2 moles of $Na_4P_2O_7$ per gram-atom of magnesium in hard water will cause the magnesium to pass through a zeolite softener, but allows calcium removal as usual. The data of Quimby (101) in table 3 also show this same trend, not only for pyrophosphate but for phosphates containing more P_2O_5.

TABLE 3

Comparison of some sodium phosphates in prevention of precipitation of magnesium or calcium soap

0.175 per cent sodium coconut oil soap at 70.5°C.

	CONCENTRATION OF PHOSPHATE REQUIRED, G./100 CC.*			
	Experimental, 5 grain hardness†		Theory (M = Ca or Mg)	
	Mg	Ca	Assumed complex	Minimum concentration
$Na_4P_2O_7$ (purified)		0.81	$Na_2MP_2O_7$	0.023
$Na_4P_2O_7$ (commercial)	0.025	0.82	$Na_2MP_2O_7$	0.023
$Na_5P_3O_{10}$ (laboratory preparation)		0.54	$Na_3MP_3O_{10}$	0.030
			$NaM_2P_3O_{10}$	0.015
"$Na_6P_4O_{13}$" (commercial, crystalline)	0.022	0.35		
Glass (commercial, 65.0 per cent P_2O_5)	0.021	0.085		
$Na_3P_3O_9$ (laboratory preparation)		>1.8		
Glass = "$Na_6P_6O_{18}$" (69.6 per cent P_2O_5)			"$Na_2M_2P_6O_{18}$"	0.026
			"$M_3P_6O_{18}$"	0.017

* *Method:* Two solutions ((A) phosphate + soap and (B) hardness) are heated separately to 70.5°C. and then mixed; the turbidity is read by a photocell 10 min. later. The phosphate concentration is varied until the amount required to transmit 80 per cent of the light transmitted by distilled water is found.

† Equivalent to 85 p.p.m. $CaCO_3$.

In the last column of table 3 are some estimates of the minimum amount of phosphate that would be required to form the indicated complex. While there is no phase diagram available on the nature of the complexes, there is some basis for the complexes $Na_2MP_2O_7$ (40, 136), $Na_3MP_3O_{10}$, $NaM_2P_3O_{10}$ (11, 13, 36, 58, 117), and "$Na_2M_2P_6O_{18}$" (20, 22, 126, 128), the last being accepted merely as a convenient way of expressing the approximate behavior of the glass (69.6 per cent P_2O_5). Whatever the chemical nature of the complexes, it is evident that little or no excess sodium phosphate is required to suppress magnesium, but a large excess is needed to suppress calcium. While the amount of phosphate required for magnesium suppression is practically independent of phosphate composition, the requirement for calcium suppression decreases as the P_2O_5 content of the phosphate increases. This trend was confirmed for a series of sodium phosphates made by neutralizing H_3PO_4–P_2O_5 mixtures homogenized in sealed tubes at 180–400°C.

All sodium phosphates in this composition range show pH sensitivity (101). For example, at 54.5°C. the amount of a glass (65 per cent P_2O_5) required for calcium suppression increased threefold as the sodium phosphate pH (in 2.4 g. per 100 ml. solution) increased from 6.7 to 9.1. Hall (45) has observed a similar result for such a glass, and Rudy *et al.* (110) observed it with $Na_5P_3O_{10}$ as well as with the glass. The effect decreases as the pyrophosphate P_2O_5 content is approached, pyrophosphate itself not being much affected.

Systematic variation of P_2O_5 content in the sodium phosphates made from homogenized H_3PO_4–P_2O_5 mixtures gave the surprising result of defining a curve practically coinciding with the data of table 3 when the amount of phosphate required to suppress calcium was plotted against the P_2O_5 content of the sodium phosphate (101).

The soluble, crystalline salt $NaPO_3$ (or $Na_3P_3O_9$) is quite ineffective as a calcium suppressor, as shown by table 3. This has been observed repeatedly (46, 59, 94).

C. THRESHOLD TREATMENT

Condensed sodium phosphates, especially $Na_5P_3O_{10}$ and the glasses, act as deterrents to the crystallization of calcium carbonate in very low concentrations, e. g., 1–5 parts per million (21, 33, 53, 103, 104). Such concentrations are far below those required for calcium suppression, yet the two phenomena are probably related. The inhibiting action is presumably due to adsorption of the complex phosphate on submicroscopic nuclei (21, 53) so as to prevent growth and hence prevent precipitation. Microscopic studies at concentrations below the threshold show that the sodium phosphates cause distortion of the calcite crystals. This distortion increases and the amount of crystallization decreases as the concentration of complex phosphate increases until, at the threshold concentration, no crystallization occurs.

Crystalline $NaPO_3$ has negligible inhibiting action but becomes quite effective in the presence of alkali (21), presumably owing to conversion to $Na_5P_3O_{10}$ (see Bell (10)).

In addition to preventing the deposition of calcium carbonate scale, these low concentrations will dissolve old deposits of calcium carbonate scale if such solutions are kept flowing through the pipes for several months (104).

D. DISCUSSION

Data obtained in this field to date have not been intended to delimit the homogeneous-solution field in a five-component system, such as Na_2O–CaO–P_2O_5–fatty acid–H_2O, but rather as a crude means of detecting differences between complex phosphates made in various ways. Nevertheless, by a suitable modification of the techniques tested one should be able to map out this homogeneous area provided pH were properly controlled. Such a study would add greatly to our understanding of what takes place in the prevention of precipitation of calcium soap, particularly if the work extended far enough over into the heterogeneous area to reveal the composition of the saturating solids.

V. CHARACTERIZATION AND ANALYSIS

Before considering ordinary chemical methods of analysis, attention will be given to some special methods of characterization which tell more about the fundamental nature of the soluble phosphates both as crystals (or glass) and in solutions.

A. ACID DISSOCIATION CONSTANTS

Quite a variety of dissociation constants can be found for H_3PO_4 and $H_4P_2O_7$ (2, 7, 51, 61, 63, 68, 69, 85, 86, 89, 129, 133). For H_3PO_4 the values of Harned and Owen (51) are 7.516×10^{-3} and 6.226×10^{-8} for the first and second hydrogens, respectively, at 25°C. The value for the third hydrogen is probably near 2×10^{-13} (63). For $H_4P_2O_7$ at 18°C. Abbott and Bray (28), whose values for H_3PO_4 are higher than those of Harned and Owen by a factor of 1.5 to 3, reported 1.4×10^{-1}, 1.1×10^{-2}, 2.9×10^{-7}, and 3.6×10^{-9} for the four hydrogens. No data are available for $Na_5P_3O_{10}$ except electrometric titration curves, which show that the first three hydrogens are strong (as are the first two in $H_4P_2O_7$), the fourth is comparable to the third in $H_4P_2O_7$, and the fifth is slightly stronger than the fourth in $H_4P_2O_7$. Treadwell and Leutwyler (129) report that the weakest hydrogen in "hexametaphosphoric acid" has a dissociation constant of 1.6×10^{-2}, i. e., all hydrogens are strong.

B. MOLECULAR WEIGHT

1. Freezing-point lowering in water

If one uses the freezing-point lowering of water for determination of the molecular weights of salts containing polyvalent ions, he must either go to very high dilutions where thermal readings to 0.0001°C. will be needed, or he must make use of the interionic attraction theory at moderately low concentrations. Either method is tedious and exacting, with the result that very few satisfactory determinations of molecular weight have been made using water as the solvent for polyvalent salts. Nylén (91) has applied the latter variation to the common, soluble, crystalline $NaPO_3$, which is thermodynamically stable in the Na_2O–P_2O_5 system at elevated temperatures below the melting point, 620–625°C. He concluded that the salt is the trimer $Na_3P_3O_9$.

2. Eutectic or transition lowering

Muller (87) has developed a much better method, involving either the lowering of some eutectic point or the lowering of some transition point. The plot of molal lowering *vs.* concentration is much less curved, in fact, it is often nearly a straight line even at relatively high concentrations, thus making simpler, cruder technique applicable. Thus for $Na_4P_2O_7$ the KNO_3–ice eutectic gives a line extrapolating back to 5 molal lowerings; $K_4P_2O_7$ would give a line extrapolating to 1 molal lowering, since ions common to both the added salt and the eutectic have no effect in lowering the eutectic point. In the same way one can use the transition point of salt hydrates, e. g.:

$$Na_2SO_4 \cdot 10H_2O \xrightleftharpoons{32.4°C.} Na_2SO_4 + 10H_2O$$

$$CaCl_2 \cdot 6H_2O \xrightleftharpoons{30.2°C.} CaCl_2 \cdot 4H_2O + 2H_2O$$

$Na_4P_2O_7$ would give 1 molal lowering of the $Na_2SO_4 \cdot 10H_2O$ transition, and 5 molal lowerings of the $CaCl_2 \cdot 6H_2O$ transition.

Bonneman-Bémia (13) has made much use of the $Na_2SO_4 \cdot 10H_2O$ transition. For example, purified $Na_5P_3O_{10}$ gave a molecular weight of 372 (theory 368), while metaphosphate made by heating NaH_2PO_4 at temperatures in the interval 300–600°C. gave 307–317 (theory 306 for $Na_3P_3O_9$). But the most interesting use involved dehydration of $Na_5P_3O_{10} \cdot 6H_2O$ in an oven at various temperatures. The molecular weight, calculated to the anhydrous basis, remained constant at nearly the theoretical 368 as the temperature of heating for half-hour periods was raised from 85° to 100°C., but it dropped sharply to a minimum of 245 near 120°C. and then increased gradually back to 368 at temperatures near 300°C. Longer heating, at lower temperatures, e. g., 12 hr. at 150°C., restored the original molecular weight of 368, but heating at 120°C. failed to do so. Bonneman-Bémia postulated that the triphosphate decomposed as follows:

$$2Na_5P_3O_{10} \cdot 6H_2O = 2Na_4P_2O_7 + Na_2P_2O_6$$

This hypothesis ignores the fact that not all of the water was removed in these experiments, and will not explain the pH and acid–base titration data to be presented later.

3. *Sedimentation of* Na_2O–P_2O_5–H_2O *glasses*

Lamm (69a) has applied the equilibrium ultracentrifuge to glasses made by quenching 0.5-g. charges of melt as follows: (*a*) NaH_2PO_4 heated to 600–650°C. for a short time or (*b*) $Na_3P_3O_9$ heated for 3 hr. at 650°C. For the former Lamm[9] calculated a molecular weight of about 13,000, for the latter a value of 12,600. Such values are equivalent to about 125 monomeric $NaPO_3$ units. While such data should reveal the size of the sedimenting particle under the particular conditions used, they need not necessarily mean that all 125 units are linked by P—O valence bonds. A micelle made up, say, of twenty-one hexameric units associated into some sort of a complex would lead to such a result. There is perhaps less evident reason for such "water-soluble" polyphosphate chains to associate than there is for "water-insoluble" hydrocarbon chains in solutions of paraffin chain salts; nevertheless, the evidence available does not exclude the possibility of a micelle. If a micelle forms, its size is probably a function of electrolyte concentration.

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[9] Tests made in 0.25–0.40 per cent solution of the phosphate glass in 0.1 *N* sodium thiocyanate at 11°C.

Lamm states that these glasses are remarkably monodisperse, differing from synthetic organic polymers, which reveal a rather wide range of molecular sizes in the equilibrium centrifuge. He states also that the particle size decreases with time of heating the melt from NaH_2PO_4.

The significance of these sedimentation data will be more apparent when other methods of determining particle size have been applied. For example, molecular-weight determinations by transition-point lowering with both $Na_2SO_4 \cdot 10H_2O$ and $CaCl_2 \cdot 6H_2O$ would give an idea of what proportion of the sodium ions are free. However, it may be mentioned that evidence is gathering to show that these Na_2O–P_2O_5–H_2O glasses have a high molecular weight. Karbe and Jander (64) interpreted dialysis data as indicating a molecular weight of about 4000 for glass from NaH_2PO_4 heated to 650°C. Samuelson (112a) from titration of acid end groups (see section on acid–base titration) calculated a value of 10,800 for such a product.

C. ELECTRICAL CONDUCTIVITY

The Oswald–Walden–Bredig rule says that, if one determines the equivalent conductance at N/32 and N/1024, the difference divided by 10 gives approximately the product of the ionic valences, e.g., 2 for $MgCl_2$ and 4 for $MgSO_4$ or $Na_4P_2O_7$. Application of this rule to $Na_5P_3O_{10}$ gave 4.9 (13), confirming the pentavalence of the P_3O_{10} ion and the ionization of all five sodium ions. For the common, soluble, crystalline $NaPO_3$ this quantity is 3.0 (97, 134), confirming the trimeric formula $Na_3P_3O_9$.

For glass of approximately the meta P_2O_5 content, e.g., the product of rapidly chilling a melt made from NaH_2PO_4, $Na_2H_2P_2O_7$, or $Na_3P_3O_9$, application of this rule has given values from 2 to 4 (96, 111, 134). On the basis of an assumed hexameric formula, Pascal *et al.* (96, 97) suggested that these glasses are made of complexes such as $Na_2(Na_4P_6O_{18})$, from which only two sodium ions would separate in aqueous solution. As seen under the subject of hardness suppression, such interpretations do not accord with all the facts. How such conductivity data would be interpreted in terms of such high molecular sizes as reported by Lamm (69a) is not clear at present.

Again it is evident that more data are needed. Precise transference plus conductivity data would be very helpful in revealing the manner of ionization of these phosphate glasses.

D. X-RAY PATTERNS

X-rays have been used on sodium phosphates chiefly for identification purposes. However, the structures of the ortho- and pyrophosphate ions have been determined by crystal analysis on less soluble salts.

1. Identification

X-ray patterns are already available in the Dow system (26, 50) for most of the twenty-seven known crystalline sodium phosphates, but some of the patterns refer to substances of doubtful purity. In favorable cases, however, a moderately pure crystalline sodium phosphate can be identified by x-rays alone. Furthermore, the presence of two or more sodium phosphates in a given product has been demonstrated in the phase studies previously described (3, 13, 16, 17, 42, 57, 95, 115).

It should also be possible not only to identify, but also to estimate roughly the concentration of each component present in a complex mixture of crystalline sodium phosphates. However, this requires considerable time and study beforehand with known mixtures to learn how to interpret the complex x-ray pattern which results. Such analysis is often further complicated because the phosphates are mixed with other substances such as silicates, carbonates, soap, or synthetic detergents. Even when x-ray analysis proves inadequate in a particular application, it may prove valuable as a supplement to chemical analyses.

2. Ionic structure

Complete x-ray analysis of phosphate structure has been carried out only for the ortho- (98) and the pyrophosphates (72), with structures established as indicated in table 4. The classical valence formulas do not convey the proper impression. In Na_3PO_4, for example, all four oxygens are known to be equivalent, whereas the classical formula indicates that three are mutually equivalent but different from the fourth. The Werner-type formula is more correct in this respect but does not indicate the structure. The PO_4^{---} ion is tetrahedral, with the P—O distance at 1.54 A. (98).

Studies of crystal structure by x-ray have shown that the pyrophosphate ion does have the oxygen bridge between the two phosphorus atoms (72), as indicated by the classical formula. Here again the classical formula is inadequate, because the six oxygen atoms not in the bridge are equivalent, not divided into two groups (4 and 2). The deficiency of the Werner-type formula is more apparent here than it was for the orthophosphate ion.

By analogy with the pyrophosphate we may represent the triphosphate structure in the same manner by simply adding a second oxygen bridge.

Some investigators have suggested that $Na_3P_3O_9$ is cyclic (101[10], 110, 129), involving three oxygen bridges, but so far there is no proof beyond the molecular-weight and conductivity data already mentioned.

E. OPTICAL DATA

1. Properties of pure compounds

Ingerson and Morey (60) have recently collected all available optical data on sodium phosphates (8, 12, 28, 65, 95, 114) and added a few of their own. Their results are collected in table 5, together with a few observations made in this laboratory[11] (101).

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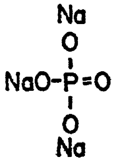
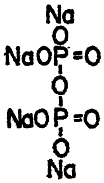
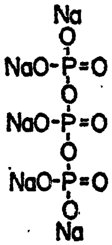
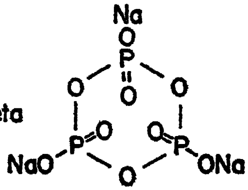
[10] Suggested to the writer by Dr. A. S. Richardson during the course of our studies.

The "$Na_3PO_4 \cdot 12H_2O$" is probably the variable hydrate $Na_3PO_4 \cdot xNaOH \cdot (12 - x)H_2O$. The salt $Na_3PO_4 \cdot 7H_2O$ has usually been given as the $8H_2O$ or $6H_2O$ compound, but Hall (44) claimed to have made the heptahydrate by crystallization of the tertiary salt from 50 per cent sodium hydroxide at 77°C If the interpretation given earlier for hydrates of Na_3PO_4 is correct, this salt is probably $Na_3PO_4 \cdot 6H_2O$, since Ingerson and Morey (60) crystallized it by evaporation of its aqueous solution at 83°C. The lower hydrate is given as the monohydrate by several authors (e. g., 60, 114), but as the hemihydrate by others (e. g., 82). More work is obviously needed on hydrates of Na_3PO_4.

TABLE 4

Structural formulae of phosphates

	Classical	Werner	Crystal Structure
Ortho		$Na_3^+[PO_4]^{\equiv}$	O O P O O All O alike P-O = 1.54 A.
Pyro		$Na_4^+[P_2O_7]^{\equiv}$	O O O P O' P O O O Six O alike one different P-O = 1.56 A. P-O' = 1.52 A.
Tri		$Na_5^+[P_3O_{10}]^{\equiv=}$	
Trimeta		$Na_3^+[P_3O_9]^{\equiv}$	

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[11] The writer is indebted to Dr. F. B. Rosevear of this laboratory for encouragement and assistance in the microscopic study of these compounds.

Our own studies (101) have made it clear that the data of Dufet (28), quoted in table 5, for $Na_2HPO_4 \cdot 12H_2O$ refer to the low-temperature or β-modification.

TABLE 5

Optical properties of sodium phosphates

FORMULA	OPTIC ANGLE	OPTIC SIGN	α	ω	β	ε	γ	REFERENCE
$Na_2O:P_2O_5 = 3:1$								
	degrees							
"$Na_3PO_4 \cdot 12H_2O$"		+		1.4458		1.4524		(28)
"$Na_3PO_4 \cdot 7H_2O$" *	85	+	1.462		1.470		1.478	(60)
"$Na_3PO_4 \cdot H_2O$"		+		1.497		1.522		(60, 114)
Na_3PO_4	76	+	1.493		1.499		1.508	(114)
$Na_2O:P_2O_5 = 2:1$								
$Na_2HPO_4 \cdot 12H_2O$	57	−	1.432		1.436		1.437	(28, 101)
$Na_2HPO_4 \cdot 7H_2O$	39	+	1.4412		1.4424		1.4526	(28)
$Na_2HPO_4 \cdot 2H_2O$	*ca.* 80	+	1.450		1.461		1.477	(60)
Na_2HPO_4	78	+	1.483		1.499		1.525	(60)
$Na_4P_2O_7 \cdot 10H_2O$	60.5	+	1.4499		1.4525		1.4604	(28)
$Na_4P_2O_7$	40	+	1.475		1.477		1.496	(60, 95)
$Na_2O:P_2O_5 = 5:3$								
$Na_5P_3O_{10} \cdot 6H_2O$†	20	+	1.449		1.450		1.482	(12, 60)
$Na_5P_3O_{10}$ I	21	+	1.477		1.478		1.504	(60, 95)
$Na_5P_3O_{10}$ II	57	+	1.470		1.477		1.502	(60, 95)
$Na_2O:P_2O_5 = 1:1$								
$NaH_2PO_4 \cdot 2H_2O$	82	−	1.4400		1.4625		1.4818	(28)
$NaH_2PO_4 \cdot H_2O$	29	−	1.4557		1.4852		1.4873	(28)
NaH_2PO_4	64	−	1.481		1.507		1.517	(60)
$Na_2H_2P_2O_7 \cdot 6H_2O$	32	−	1.4599		1.4645		1.4649	(28)
$Na_2H_2P_2O_7$‡			1.510				1.517	(60)
$Na_3P_3O_9 \cdot 6H_2O$*	77 ± 3	−	1.432				1.441	(60, 65, 101)
$Na_3P_3O_9 = NaPO_3$ I	80	−	1.474		1.478		1.480	(8, 60, 95)
$NaPO_3$ II*	78	+	1.498		1.510		1.529	(60, 95)

* See text.

† Common view of rectangular plate gives $\alpha' = 1.4495$ and $\gamma' = 1.4700 \pm 0.001$ (10 A).

‡ Fine needles, positive elongation; indices really α' and γ', since orientation of needles could not be checked.

The α-form was encountered as β-pseudomorphs when β-crystals were warmed up to 31–32°C. on the microscope stage. Crystals of α-$Na_2HPO_4 \cdot 12H_2O$ (e. g., crystals formed directly from aqueous solutions at 30–35°C.) have not been studied microscopically, but Hammick *et al.* (49) stated that these crystals look like the monoclinic β-crystals.

As suggested in Mellor's treatise (79), the soluble sodium metaphosphate usually crystallizes in "triclinic rhombohedra" of $NaPO_3 \cdot 2H_2O$ (or $Na_3P_3O_9 \cdot 6H_2O$). The acute bisectrix interference figure reveals the absence of planes or axes of symmetry and therefore confirms the triclinic nature of this crystal (101). Like Ingerson and Morey we encountered what appeared to be a metastable form,[12] for it was encountered, if at all, only in early stages of fast crystallizations and always disappeared later (if left in contact with the mother liquor), leaving exclusively the triclinic rhombohedra already described.

This "metastable" form is probably the anhydrous salt, as suggested by Ingerson and Morey.

TABLE 6

Microscopic identification of phosphate-type crystals from aqueous solution of pH 9–10

PHOSPHATE		CHARACTERISTIC PROPERTIES
Type	Formula	
Ortho...........	$Na_2HPO_4 \cdot 12H_2O$ (β)	Diamonds; 60° silhouette angle; symmetrical extinction
Pyro...........	$Na_4P_2O_7 \cdot 10H_2O$	Columnar to acicular; parallel extinction; negative elongation
Tri.............	$Na_5P_3O_{10} \cdot 6H_2O$	Rectangular plates; oblique extinction 8–9°; often twinned with extinction 8–9° to twinning plane
Trimeta.........	$Na_3P_3O_3 \cdot 6H_2O$	Rhombohedra; oblique extinction; 70–75° silhouette angle

2. Microscopic identification

Given a sample known to be an Na_2O–P_2O_5–H_2O preparation, the phosphate type among the crystallizable species can be identified microscopically rather simply. The aqueous solution is adjusted to pH 9–10, a drop is allowed to evaporate on a microscope slide below 30°C., and the crystals are examined. Fairly pure compounds reveal the characteristics listed in table 6. Phosphate mixtures give less well defined crystals, but these characteristics are still useful. The presence of Na_2O–P_2O_5–H_2O glass sometimes alters these forms considerably. For example, a rectangular plate of $Na_5P_3O_{10} \cdot 6H_2O$ will be smaller and twins more complex (i.e., appear to be made of four or six instead of two parts), but the 8–9° extinction is still observable. Purification by one or two quick recrystallizations[13] removes enough of the glass to give more simple crystals. Presence of non-phosphatic material may render such simple tests unreliable. In this case successful microscopic identification is still possible by use of the quantitative data in table 5, which may be applied without pH adjustment provided the unknown does not happen to be a compound for which the optical data are still lacking.

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[12] Near optic axis view was often obtained showing symmetrical extinction, positive, large optic angle, $r > v$, $\gamma' = 1.507$, $\alpha' = 1.495$.

[13] It is very convenient to add alcohol to induce rapid crystallization.

F. CHEMICAL ANALYSIS

From the foregoing discussion, it is evident that there are more than four chemical or ionic types of phosphate to be determined. The glasses act in some respects so differently from the crystallized preparations having the same P_2O_5 content that it often proves useful to estimate this type of phosphate separately. While a mixture of ortho-, pyro-, and metaphosphates can be correctly analyzed by known and reliable chemical methods, the addition of triphosphate and/or phosphate glass greatly complicates the analysis. At present no thoroughly reliable method of analysis is available for complex sodium phosphate mixtures.

Attempts have been made to use prevention of the precipitation of barium ion or calcium ion as the basis of an analytical method (4, 22, 39), but they have not proven successful (38). This is to be expected because of the sensitivity to temperature and pH which leads to a large experimental error. The more reliable chemical methods called upon so far have involved oven-drying, ignition, hydrolysis, precipitation, and acid–base titrations.

1. Oven drying

Many of the salts discussed lose all water of crystallization when dried in an oven at 105°C., e. g., NaH_2PO_4, Na_2HPO_4, $Na_2H_2P_2O_7$,[14] and $Na_4P_2O_7$. Hydrates of Na_3PO_4 dry down to an indefinite water content near that for the formula $Na_3PO_4 \cdot H_2O$; for complete removal of water from "$Na_3PO_4 \cdot 12H_2O$" see the section on ignition below. Both Salzer (112) and Partington and Wallsom (93) say that $Na_3HP_2O_7 \cdot H_2O$ is stable up to 170° to 190°C., but this should be checked by modern methods.

The triphosphate $Na_5P_3O_{10} \cdot 6H_2O$ loses most of its water of crystallization (13) but suffers some hydrolysis to acid pyro- and orthophosphates, as shown in the section on acid–base titrations. The metaphosphate $Na_3P_3O_9 \cdot 6H_2O$ would be expected to behave in a similar manner, but so far proof is lacking.

2. Ignition

Ignition to red heat will dehydrate completely most of the phosphates, removing both water of hydration and water of constitution. Two sources of difficulty are known, however.

The complex salt $Na_3PO_4 \cdot xNaOH \cdot (12 - x)H_2O$ will not lose all its water without heating to a temperature of 1000°C. where it also loses sodium hydroxide (82). But if some $Na_4P_2O_7$ is added to the sample before ignition, the extra sodium oxide is retained while all water is lost, leaving the equivalent of $Na_3PO_4 \cdot xNa_2O$ combined and/or mixed with $Na_4P_2O_7$. Probably moisture in "$Na_3PO_4 \cdot$

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[14] Possibly some hydrolysis to orthophosphate occurs during the drying of $Na_2H_2P_2O_7 \cdot 6H_2O$.

$12H_2O$" or in any hydrate of Na_3PO_4 should be determined by this method regularly; otherwise there is no certainty about the amount of water retained.

The other case requiring caution is any ignition in which the sample melts (especially tri- and metaphosphate samples) at high temperatures (550–900°C.). The last trace of water of constitution is lost extremely slowly from such melts. For example, most preparations of so-called "glassy meta" show an acid reaction, which means incomplete removal of the water of constitution during preparation. Re-ignition of such a glass in the usual way does not remove all water of constitution (101). Better results are obtained by grinding such a glass very fine and heating to constant weight at 500°C. For best results the glass should be mixed intimately with ten times its weight of $Na_4P_2O_7$ and ignited in the usual way.

In short, we can say that best results are obtained by insuring that the final composition of the ignited product will approach that of $Na_4P_2O_7$.

3. Acid–base titrations

(a) Acids

Sample electrometric titration curves for aqueous solutions of phosphoric acids that are commercially available are shown in figure 5. Within the range of the glass electrode it is possible to neutralize only two of the hydrogens in H_3PO_4, one at pH 4.3 and the other at pH 9.3. $H_4P_2O_7$, however, can be completely neutralized,—two hydrogens at pH 4.3, the third at pH 7.2, and the fourth at pH 10.2. The so-called tetraphosphoric acid is a mixture of H_3PO_4, $H_4P_2O_7$, $(HPO_3)_x$, and possibly $H_5P_3O_{10}$ and other polybasic acids. The presence of both ortho- and pyrophosphoric acid can be seen from the fact that the neutralization curve has a very *long* steep portion at the upper end point.

The first (or pH 4.3) end point measures the total P_2O_5 in the acid. When any acid derivable from P_2O_5 and water is titrated to pH 4.3, one titratable H has been neutralized per atom of phosphorus. This relationship seems to be perfectly general, applying to all phosphoric acids including P_2O_5 itself (37, 76, 101). It therefore gives at once the total P_2O_5 content.

Since metaphosphoric acid is already completely neutralized at pH 4.3, the spread between the 4.3 and the upper end point depends in general on ortho-, pyro-, and triphosphates.[15] If we assume with Gerber and Miles (37) that liquid acids in the ortho-to-meta range contain no triphosphoric acid or higher polybasic acids, then this spread gives the sum of the ortho- and pyrophosphoric acids, for each contributes one titratable hydrogen per atom of phosphorus.

By adding another titration designed to give the amount of orthophosphoric acid, Gerber and Miles were able to express the composition of condensed phosphoric acids in terms of the relative amounts of ortho-, pyro-, and metaphosphoric acids (triphosphoric acid assumed absent). The titration of the orthophosphoric acid involved adding an excess of silver nitrate after having titrated to the end point near pH 10; only orthophosphoric acid liberates titratable hydrogen as the phosphates are precipitated by silver, giving one hydrogen for each orthophosphoric acid phosphorus atom.[16]

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[15] This assumes the absence of higher polyphosphoric acids, i.e., $H_{n+2}P_nO_{3n+1}$ with $n > 3$. While there is no evidence at present for the existence of such acids, they may be found in liquid phosphoric acids containing 75 per cent or more of P_2O_5.

This type of analysis led Gerber and Miles to the composition diagram of figure 6. Except for the terminal compositions H_3PO_4 and HPO_3, all appear as mixtures of ortho, pyro, and meta acids. The form of the curve defined by the points suggests an equilibrium (homogeneous) of the type indicated by the equation:

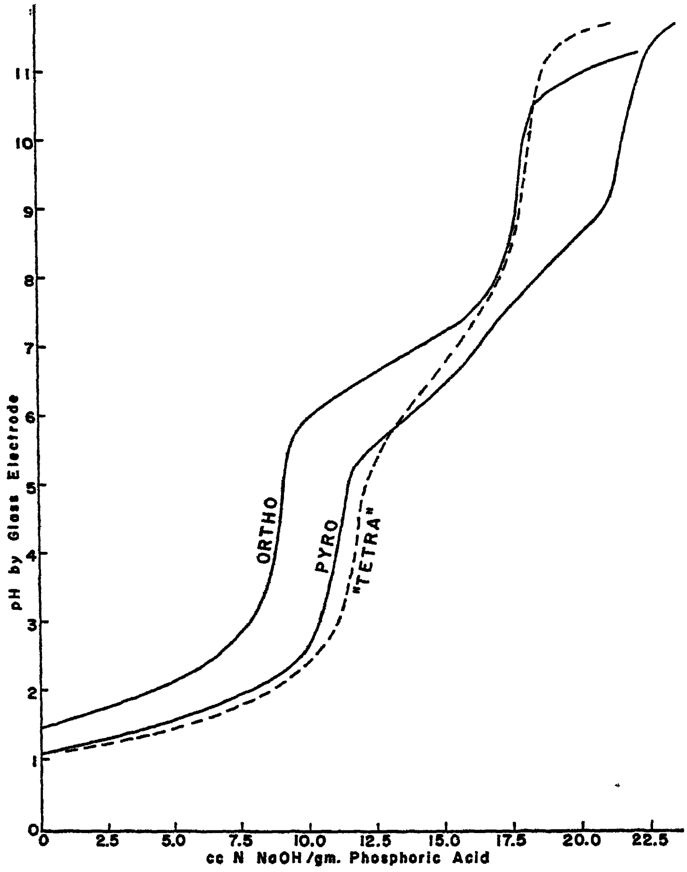

FIG. 5. Electrometric titration of some phosphoric acids

$$\underset{\text{Pyro}}{2P_2O_5 \cdot 2H_2O} \rightleftharpoons \underset{\text{Ortho}}{P_2O_5 \cdot 3H_2O} + \underset{\text{Meta}}{P_2O_5 \cdot H_2O}$$

The smooth curve of figure 6 is given by $(P_2O_5 \cdot 3H_2O)(P_2O_5 \cdot H_2O)/(P_2O_5 \cdot 2H_2O)^2 = K = 0.068$ for concentrations in weight per cent. While this equilibrium may not be expressed in the proper terms, it is probable that some such equilibrium will be substantiated by further investigation. $H_5P_3O_{10}$ and higher polybasic acids may be found to exist in such mixtures and other degrees of polymerization, especially for metaphosphoric acid, may be involved.

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[16] Treadwell and Leutwyler (129) have used the same procedure for determining ortho P_2O_5, except that Ba^{++} was used to precipitate the phosphates at pH 9.6. Germain (39) used Ca^{++} to precipitate the phosphates, but the results were not quantitative.

An appreciation of the significance of the above chemical equilibrium can be obtained by considering an experiment made by Gerber and Miles and checked in our laboratory. Pyrophosphoric acid can be obtained as a crystalline solid and is so sold by chemical distributors. If this acid is dissolved in cold water and analyzed by the titration method of Gerber and Miles, it will be found to contain 95–100 per cent $H_4P_2O_7$. Now if this crystalline acid is heated above its melting point (*ca.* 61°C.) to liquefy it and is restored to room temperature, it remains liquid for a long time (weeks or months) if no $H_4P_2O_7$ seed is added or it is not stirred. A sample of this liquid tested by the Gerber and Miles method yields the composition 65 per cent $H_4P_2O_7$ plus 35 per cent of an equimolar mixture of H_3PO_4 and HPO_3. Evidently when heated above their melting points these acids decompose in accord with some such chemical equilibrium as indicated by the equation above. The true equilibrium is not readily restored by cooling back below the melting point, owing to the very great viscosity of these acids. Thus Gerber and Miles really measured some approach to the chemical equilibrium which obtains at some temperature above the melting point.

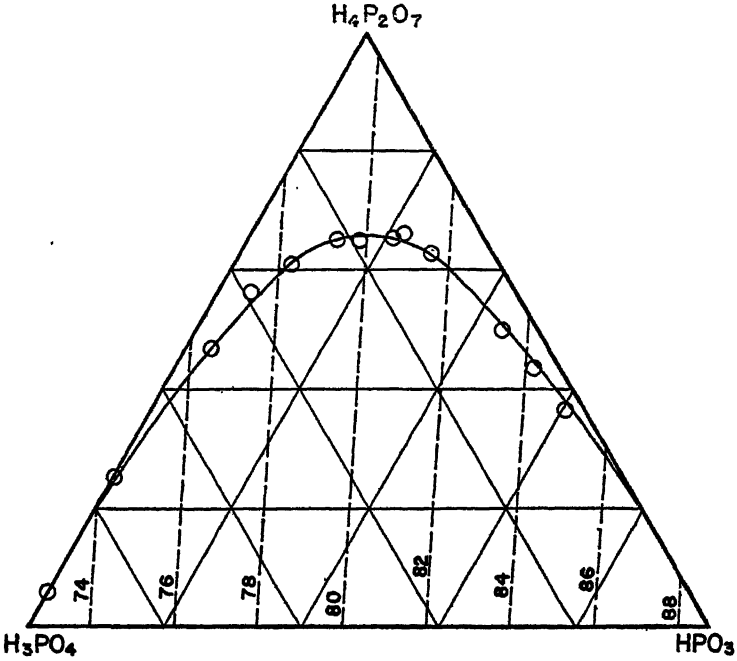

FIG. 6. Compositions of anhydrous phosphoric acids. ———, compositions given by equilibrium constant; O, compositions found by colorimetric pH titration;— — —, lines of equal per cent P_2O_5.

Such data make it very clear why the H_2O–P_2O_5 binary system is so poorly understood. It is probable that true equilibria could be approached by cooling to some low temperature to induce seeding and then restoring to a temperature a few degrees below the stable melting point, as was done by Schierloh (113) for glycerol.

(b) Salts

Acid–base titrations can be used just as effectively on phosphate salts, except that they cannot reveal the P_2O_5 content. Sample curves are given for condensed sodium phosphates by Schwartz and Munter (116). The curve for $Na_4P_2O_7$ shows exactly the same characteristics as the curve for the forward titration of $H_4P_2O_7$ in figure 5—namely, prominent waves at pH 4 and pH 10 and a slight wave at 7. The spread between corresponding end points agrees quantitatively on an equal P_2O_5 basis.

TABLE 7

Chemical equivalents involved in phosphate titrations before and after hydrolysis

PHOSPHATE TYPE*	SPREAD BETWEEN UPPER AND LOWER END POINTS, EQUIVALENTS PER PHOSPHORUS ATOM			SPREAD BETWEEN END POINTS AS MILLILITERS OF *N* REAGENT PER G. P_2O_5		
	Before	After	Difference	Before	After	Difference
Ortho	1	1	0	14.08	14.08	0.00
Pyro	1	1	0	14.08	14.08	0.00
Tri	$\frac{2}{3}$	1	$\frac{1}{3}$	9.39	14.08	4.69
Meta	0	1	1	0.00	14.08	14.08

* Refers to compounds, i.e., crystalline species. Glasses can behave like tri- or metaphosphate, but more generally will act like a mixture of two to four of the above phosphates.

A similar curve for $Na_5P_3O_{10}$ shows that its upper end point is slightly below that of $Na_4P_2O_7$; it shows two other inflections, one at a pH of about 7 and the other at about 4, corresponding to $Na_4HP_3O_{10}$ and $Na_3H_2P_3O_{10}$, respectively (*cf.* Rudy and Schloesser (109)). Taken with the pyrophosphate curve, which also shows a slight inflection near pH 7 corresponding to $Na_3HP_2O_7$, this curve tells us that $Na_4HP_3O_{10}$ exists. Huber (58) has already crystallized one hydrate of it ($Na_4HP_3O_{10} \cdot 1.5H_2O$). The curve tells us further that the third, fourth, and fifth hydrogens of $H_5P_3O_{10}$ are comparable in strength to the second, third, and fourth hydrogens of $H_4P_2O_7$, respectively.

Electrometric titrations of condensed phosphates made so as to define the full course of the curve between pH 2.5 and 11 are very useful in estimating the condition of the product in solution. They are applicable both to crystalline compounds and to glasses (*cf.* Schwartz and Munter). The chemical relationships are summarized in table 7.

The product of heating $Na_5P_3O_{10} \cdot 6H_2O$ for 5.5 hr. at 120°C. (101) is an excellent illustration of the utility of the electrometric titration curve. The pertinent data are given in table 8. These data illustrate the need for titrating the sample both ways from its existing condition. Thus on the anhydrous basis, titration to the lower end point gives nearly the same result for $Na_5P_3O_{10}\cdot 6H_2O$ and its decomposition product. But the *spread* between the two end points is very different. Both the spread and the pH warn of the radical alteration of the $Na_5P_3O_{10}$. As stated in the footnote to table 8, the presence of orthosphosphate is indicated. The remainder is largely pyrophosphate.

Utilizing these data and the total P_2O_5 content, one can infer that the decomposition proceeded approximately as follows:

$$32Na_5P_3O_{10}\cdot 6H_2O \xrightarrow{-166H_2O}$$

$$13Na_4P_2O_7 + 26Na_3HP_2O_7 + 10NaH_2PO_4 + 8Na_2HPO_4$$

TABLE 8

Decomposition of $Na_5P_3O_{10}\cdot 6H_2O$ *in oven drying at 120°C.*

	$Na_5P_3O_{10}\cdot 6H_2O$		DECOMPOSITION PRODUCTS	
	As is	Anhydrous	As is	Anhydrous
pH of 1 g./100 cc. solution	9.8		8.0	
Milliliters of *N* alkali to pH 10 end point per gram of phosphate	0.03−	0.03	2.45	2.55
Milliliters of *N* acid to pH 4 end point per gram of phosphate	3.91+	5.08	5.12	5.34
Spread between pH 4 and pH 10 phosphate	3.94	5.11	7.57	7.89
Per cent loss on ignition	22.91*		4.01†	
Per cent total P_2O_5	45.42*	58.9	55.59	72.1
Per cent CO_2 by evolution	0.01		None	

* Theory for $Na_5P_3O_{10}\cdot 6H_2O$ is 22.70 per cent H_2O and 44.7 per cent P_2O_5.

† About 0.87 mole $H_2O/Na_5P_3O_{10}$ remains; the titration from pH 8 to pH 10 is equivalent to 2.14 per cent H_2O. If the difference between 4.01 and 2.14 is water of constitution not titrated, i.e., ortho, then there is 10.4 per cent P_2O_5 present in the ortho condition.

The mixture on the right has $\frac{3}{16}$ of its total P_2O_5 in the ortho condition ($55.6 \times 3/16 = 10.4$ per cent P_2O_5). This mixture will give a pH of about 8 in dilute solution and will lose 4.0 per cent H_2O upon ignition; it has a molecular weight of 213, a value appreciably below the 245 minimum obtained by Bonneman–Bémia (13). It is probable that more conversion to orthophosphate occurred in the writer's experiments, for Bonneman–Bémia's sample contained less water of constitution; this would account for the difference in molecular weight. The indicated type of decomposition to pyro- and orthophosphate is more in line with all the facts than Bonneman–Bémia's hypothesis that pyro- and dimetaphosphates are formed.

While their interpretation is attended by more uncertainty, acid–base titration curves are quite as useful for characterizing the Na_2O–P_2O_5–H_2O glasses. The presence of hydrogens titratable between pH 4 and pH 10 (101, 116, 128) reveals at once that a glass is not simply a metaphosphate $(NaPO_3)_n$.[17] In some cases glasses might conceivably be mixtures of $(NaPO_3)_n$ with acid salts such as NaH_2PO_4, $Na_2H_2P_2O_7$, etc.,[18] but a more likely picture is that of a polyphosphate $(Na,H)_{n+2}P_nO_{3n+1}$ plus small amounts of low-molecular-weight acid salts. When quantitative chemical methods of determining the ortho, pyro, and tri P_2O_5 content of such glasses become available, the part of the titration (pH 4 to 10) due to low-molecular salts can be calculated and deducted from the total. The remainder must be attributed to higher polyphosphates.

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[17] This follows from the fact that any phosphoric acid titrated to pH 4 with sodium hydroxide contains one sodium per phosphorus atom; hence $(NaPO_3)_n$ should contain no hydrogen titratable between pH 4 and pH 10.

On the assumption that glasses made from NaH_2PO_4 can be written $Na_nH_2P_nO_{3n+1}$, Samuelson (112a) has utilized the titration to the phenolphthalein end point to calculate the average molecular weight. The values so calculated varied from 10,800 for a fusion temperature of 650°C. to 17,200 for 950°C. Assuming no ring formation and no loss of P_2O_5 in the fusion process, Samuelson's reasoning should be correct. However, neither ring closure nor loss of P_2O_5 can be regarded as impossible at fusion temperatures. The former would lower the number of titratable hydrogens (by water loss); the latter would increase the number of titratable hydrogens per molecule. Incidentally, flow birefringence studies would be helpful in testing Samuelson's chain model for such glasses.

One use can be made of the titration of any sodium phosphate to the end point near pH 10 which is subject to no uncertainty of interpretation. Knowing the amount of orthophosphate present and the magnitude of this titration, one can calculate the amount of bound water, i.e., water of constitution. If the orthophosphate content is unknown, calculation of the titration to its water equivalent sets a lower limit to the bound-water content. Thus a glass of 66.8 per cent P_2O_5 content was found to contain at least 1.12 per cent combined water. Ignition of this sample above its melting point caused a loss of but 0.71 per cent, illustrating why the final ignition composition should be near that of $Na_4P_2O_7$ for complete water loss. Had the "pyrophosphate" ignition method been applied to the above glass and given a loss greater than 1.12 per cent, the excess could be calculated to orthophosphate at the rate of one PO_4 per hydrogen.

An acid-base titration for determining pyrophosphate was introduced by Britzke and Dragunov (18). Like the Gerber and Miles ortho determination, it is really a combination of precipitation and titration, but it is carried out volumetrically. The pH of the phosphate solution is adjusted to a point near the lower end point, e.g., to pH 3.8, and excess of zinc sulfate is added to precipitate the pyrophosphate as $Zn_2P_2O_7$ and liberate one titratable hydrogen for each pyro phosphorus atom, thus:

$$H_2P_2O_7^{--} + 2ZnSO_4 \rightarrow Zn_2P_2O_7 + 2HSO_4^-$$

The solution becomes more acid and the strong acid is titrated by restoring the pH to 3.8. Britzke and Dragunov were aware that the method was not satisfactory in the presence of much metaphosphate (see also 76), but orthophosphate does not interfere. It has since been shown that the method is not at all reliable in the presence of triphosphate (9, 62, 76, 101), but Bell has recently developed a promising modification which will be discussed under precipitation methods. It is likely that the polyphosphate molecules in Na_2O–P_2O_5–H_2O glasses also interfere.

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[18] Assuming only four chemical species present—namely, ortho, pyro, tri, and meta—one can calculate the amount of each from four data: initial pH, electrometric titration curve, true ignition loss, and total P_2O_5. Such a result may be far from the truth however, for the calculation makes no allowance for the distribution of titratable hydrogen, P_2O_5, etc., between the simple molecules (ortho, pyro, tri) and the higher polyphosphate molecules probably present in the glass.

4. Hydrolysis

In strongly acid solutions, especially at temperatures near the boiling point of water, all of the condensed phosphates are rapidly hydrated to the ortho condition. This is the basis of the well-known method of determining total P_2O_5 by the precipitation methods to be described below. It is the basis of another hydrolysis method for "apparent meta" P_2O_5.[19] In this case the phosphate mixture is titrated to the upper end point, which is near pH 9 if sufficient sodium nitrate or sodium chloride is added. Here ortho exists as Na_2HPO_4, pyro as $Na_4P_2O_7$, tri as $Na_5P_3O_{10}$, and meta as $(NaPO_3)_x$. A known amount of standard acid is added and the solution boiled to convert all phosphates to the ortho condition. The solution is then back-titrated to the original pH and the known equivalent of hydrolyzing acid deducted. The remainder measures the newly created, titratable hydrogen derived from meta- and triphosphates as indicated in table 7. Meta develops one titratable hydrogen per phosphorus atom, while tri develops $\frac{1}{3}$ of a hydrogen per phosphorus atom. Calculating all the newly created hydrogen to metaphosphate is equivalent to calling $Na_5P_3O_{10}$ a mixture of $Na_4P_2O_7$ and $NaPO_3$; hence the name "apparent meta." This method is sound chemically, but its importance is somewhat diminished by the fact that the same information can be obtained in another way. Note that the spread between the upper and lower end points is 1 H/P for all forms after hydrolysis but was 1, 1, $\frac{2}{3}$, and 0 for ortho, pyro, tri, and meta, respectively, prior to hydrolysis. Thus the difference between these two titrations is precisely equivalent to the "apparent meta" titration. Now consider three phosphate compositions having the same P_2O_5 content, say 60.4, corresponding to the hypothetical $Na_6P_4O_{13}$. We might have, for example

A. 1 mole $Na_5P_3O_{10}$ + $\frac{1}{3}$ mole $Na_3P_3O_9$ (100 per cent crystalline)
B. 0.5 mole $Na_5P_3O_{10}$ + "1 mole" glass $(Na_2O)_{3.5}(P_2O_5)_{2.5}$[20] or
C. 1 mole $Na_4P_2O_7$ + "2 moles" glass $(Na_2O)(P_2O_5)$[20]

Now should the two titrations just mentioned give the same result for all three mixtures, then only one of the two methods need be used. But, should systematic investigation show that these two titrations are frequently independent,[21] then progress in analyzing mixtures containing glass might be possible by proper interpretation of these *two* titrations.

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[19] Jones (62) and Travers and Chu (127) have used modifications of the method described here.

[20] These formulae are not meant in a molecular sense, but are used merely to express the composition of the glass.

[21] This seems rather unlikely to the writer.

5. *Precipitation*

Discussion of precipitation methods will be limited to those which have proved reliable or which give promise. Thus precipitation of orthophosphate as $MgNH_4PO_4$ or as phosphomolybdate can be, and has been, used in a variety of ways. Thus, by acidifying and boiling before precipitation, all phosphorus can be precipitated to give total P_2O_5. The determination is completed according to one of the well-known colorimetric, volumetric, or gravimetric methods given in standard texts on quantitative analysis. Orthophosphate can be determined directly in a complex phosphate mixture by precipitating it as phosphomolybdate in cold acid solution. The determination is frequently completed, after filtering off the phosphomolybdate, by a colorimetric method (14, 30, 35, 123, 130), though volumetric (62) and gravimetric methods (127) have also been used.

Jones (62) experienced difficulty in precipitating pyrophosphate as $Mn_2P_2O_7$ because triphosphate tends to precipitate with the pyrophosphate, giving high pyro P_2O_5 values. The same difficulty is encountered when the precipitation is carried out with zinc (9, 62, 72, 101). In other words, the pyrophosphate and triphosphate ions are so closely related that they undergo very similar reactions. Nevertheless, Bell (9) has succeeded in making the zinc method fairly reliable by a double precipitation; the second precipitation gets rid of the triphosphate which precipitated with pyrophosphate in the first precipitation. The $Zn_2P_2O_7$ from the second precipitation is ignited and weighed.

Bell has further found that a rough measure of the triphosphate content can be obtained by a modification of the old Britzke and Dragunov titration method for pyrophosphate (18). The pH is adjusted to 3.8, a certain excess of zinc sulfate is added, the pH is restored to 3.8 with standard alkali, and the pyrophosphate precipitate is quickly filtered off for the pyrophosphate determination just described. After deducting the pyrophosphate contribution from the sodium hydroxide titration, the remainder is converted to triphosphate by means of an empirical factor established by experiment on synthetic mixtures. While this method gives only approximate results and requires close attention to details, it has the virtue of being a direct determination and should prove useful, at least with essentially crystalline products.

Jones (62) attempted a classification of "metaphosphates" based on precipitation with barium. He found that barium precipitates the "hexametaphosphate" *almost* quantitatively in cold acid solutions, where all crystalline species remain in solution. The precipitate was filtered off, converted to orthophosphate by hydrolysis, and determined in the usual way. Jones further found that the trimetaphosphate remained in solution, while all other crystalline species of phosphate are precipitated by barium in alkaline solution. Acid hydrolysis converts this filtrate to orthophosphate for determination as above.

6. *Discussion*

A test of the Jones method has been made in this laboratory[22] (table 9). For the product of heating $Na_5P_3O_{10} \cdot 6H_2O$ for 5.5 hr. at 120°C., the ortho P_2O_5 figure is very close to the 10.4 per cent calculated above from the titration plus ignition loss and most of the remainder is in the pyro condition, as previously indicated; the meta P_2O_5 figure (5.25 per cent) is suspected of being high, for the sum of the ortho, pyro, and meta P_2O_5 exceeds the total P_2O_5 by 1.33 per cent. The analyses for both the $Na_5P_3O_{10} \cdot 6H_2O$ and the commercial "tetraphosphate," which appeared wholly crystalline under the polarizing microscope, are about what would be expected for such products.

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[22] The analyses were made by Mr. R. H. Lambert under the supervision of Mr. J. T. R.

It is felt that the Jones method represents a distinct advance over earlier methods, but it is not entirely satisfactory, as Jones has realized. Double precipitation of $Mn_2P_2O_7$ would probably improve the accuracy of the Jones pyro P_2O_5 determination; this would help the tri P_2O_5, which is a difference figure, but errors in the two "meta" determinations would still leave this result rather uncertain. The trimeta P_2O_5 figure may often be too high, owing to incomplete precipitation of the other phosphates by barium in alkaline solution (62).

TABLE 9

P_2O_5 *distribution by the Jones method*

	$Na_5P_3O_{10} \cdot 6H_2O$	PRODUCT OF HEATING $Na_5P_3O_{10} \cdot 6H_2O$ FOR 5.5 HR. AT 120°C.	COMMERCIAL "TETRAPHOSPHATE" (CRYSTALLINE)
Total P_2O_5	45.42 per cent	55.59 per cent	56.15 per cent
P_2O_5 distribution by Jones's method:			
"Hexameta" P_2O_5	None	None	None
Trimeta P_2O_5	0.59	5.25	2.84
Ortho P_2O_5	Trace	10.23	3.09
Pyro P_2O_5	0.68	41.44	27.28
Tri P_2O_5 (difference)	44.15	(−1.33?)	22.94

Summarizing the chemical methods of analysis available for complex sodium phosphates, we find that reliable methods have been developed for:

1. Total P_2O_5
2. Ortho P_2O_5 (35, 37, 62)
3. Pyro P_2O_5 (9)

Approximate methods are available for

4. Tri P_2O_5 (9)
5. Meta (trimeta) P_2O_5 (62)
6. P_2O_5 in Na_2O–P_2O_5–H_2O glass ("hexameta," 62)

In addition, there is the spread between the upper and lower end points in acid-base titrations, which measures a sort of sum of ortho P_2O_5 + pyro P_2O_5 + tri P_2O_5 (see table 7) and the reciprocally related result designated as "apparent meta" P_2O_5. It is quite probable that satisfactory analyses for many purposes can be made with the aid of these or variants of these methods. However, products containing Na_2O–P_2O_5–H_2O glass require further analytical study before a generally satisfactory method of analysis can be given.

One must not neglect valuable supplements such as observations with the polarizing microscope, x-ray patterns, or molecular-weight determinations by the methods of Muller (87). Needless to say the more general problem of analyzing complex phosphates in the presence of other substances such as borates, carbonates, silicates, soap, and/or synthetic detergents has not yet been solved.

VI. STABILITY

Discussion of the phase diagrams has already shown that, in contact with water, all phosphates except orthophosphates are thermodynamically unstable. This fact alone does not indicate the "practical" stability of condensed phosphates, for many useful organic compounds are thermodynamically unstable with respect to their elements, yet do not decompose under ordinary conditions (e. g., benzene). The question of importance is: How fast do condensed phosphates revert to orthophosphates in the presence of water under practical working conditions?

Many hydrolysis studies have been made (1, 5, 6, 8, 10, 19, 27, 39, 43, 65, 75, 84, 108, 110, 127, 131), but their utility is often marred by lack of adequate analytical methods (especially in the older work) and in some cases by lack of pH control. Watzel (131) controlled the pH and found the greatest stability for pyro P_2O_5, tri P_2O_5, and the glass of meta P_2O_5 content in the pH range 9–10; in the pH range 3–7 orthophosphate develops fastest for the glass and slowest for pyrophosphate, the triphosphate being intermediate.

Recently Bell (10) has applied more modern methods of analysis to hydration studies in strongly acid or alkaline solutions. He indicates that in hot water the course of the hydrolysis is as follows:

$$\text{“}Na_6P_6O_{18}\text{”} + 3H_2O \rightarrow Na_3P_3O_9 + 3NaH_2PO_4$$
$$Na_3P_3O_9 + 3H_2O \rightarrow 3NaH_2PO_4$$
$$Na_5P_3O_{10} + H_2O \rightarrow Na_3HP_2O_7 + Na_2HPO_4$$
$$Na_2H_2P_2O_7 + H_2O \rightarrow 2NaH_2PO_4$$
$$Na_3HP_2O_7 + H_2O \rightarrow Na_2HPO_4 + NaH_2PO_4$$
$$Na_4P_2O_7 + H_2O \rightarrow 2Na_2HPO_4$$

In hydrolyzing the "hexametaphosphate", the $Na_3P_3O_9$ formed in the first step follows the course taken by pure $Na_3P_3O_9$, as shown in the second equation. Similarly $Na_5P_3O_{10}$, which breaks first into $Na_3HP_2O_7$ and Na_2HPO_4, suffers ultimate hydrolysis to Na_2HPO_4 and NaH_2PO_4, as indicated by the equation for the hydrolysis of $Na_3HP_2O_7$.

In 1 per cent sodium hydroxide solution the hydrolysis of $Na_3P_3O_9$ follows a different course:

$$Na_3P_3O_9 + 2NaOH \rightarrow Na_5P_3O_{10} + H_2O$$

Except during the rapid conversion to triphosphate, no trimetaphosphate is present; hence this hydrolysis follows the course indicated for triphosphate in water above.

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REACTIONS OF NON-ENOLIZABLE KETONES IN SUNLIGHT

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I. INTRODUCTION

During the last decade, interest in the photochemical reactions of organic compounds has centered mainly on their mechanism; the number of new photochemical reactions of synthetic value discovered during this period is very limited, in particular with regard to those reactions carried out in sunlight.

But, during the same period, considerable progress has been made in the field of non-enolizable ketones in connection with photochemical reactions in sunlight; some of them are of special interest because they furnish an easy way to substances so far unavailable by dark reactions, a fact which justifies their review in this paper.

Photochemical syntheses utilizing sunlight are not in general, and especially on an industrial scale, suitable for use in Europe or in the northern part of America, where sunlight is unreliable and often unavailable; this is also the case in many other parts of the world. In tropical and many subtropical regions, however, conditions are different, so that the prospects for the *industrial* development of such syntheses can be considered promising.

II. PHOTOCHEMICAL FORMATION OF PINACOLS FROM DIARYL KETONES IN THE PRESENCE OF ALCOHOLS, AND RELATED REACTIONS

Ciamician and Silber (10) have shown that when an ethyl alcohol solution of benzophenone is exposed to sunlight, the alcohol is oxidized to aldehyde and the ketone reduced to benzopinacol; in the place of ethyl alcohol, isopropyl alcohol (12) has frequently been used.

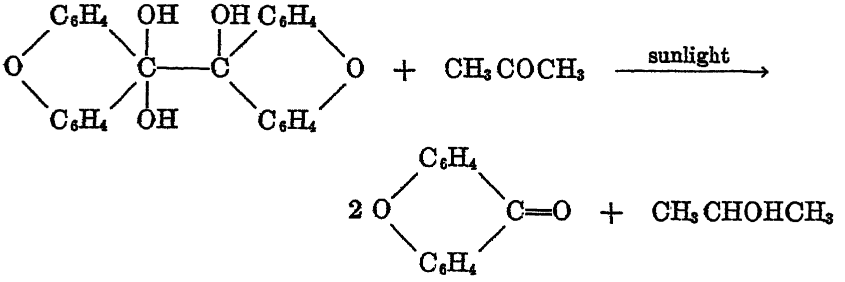
$$2C_6H_5COC_6H_5 + CH_3CHOHCH_3 \xrightarrow{\text{sunlight}} (C_6H_5)_2COH\,COH(C_6H_5)_2 + CH_3COCH_3 \quad \text{(A)}$$

A large number of benzophenone derivatives react in the presence of alcohols as above: e.g., p,p'-dimethoxybenzophenone (26, 43), p,p'-dichlorobenzophenone (43), and 4-phenylbenzophenone (3).

Banchetti (5) has investigated the action of hydrogen chloride gas on a mixture of benzophenone in isopropyl alcohol and benzene in sunlight.

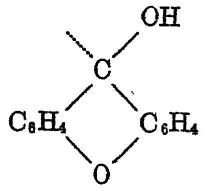
A. Aromatic ketones which are resistant to alcohols in sunlight

Bachmann (3) has reported that fluorenone is stable to isopropyl alcohol in sunlight, and Schönberg and Mustafa (43) were unable to reduce xanthone to the corresponding pinacol. They found that in the case of xanthopinacol and fluorenopinacol, the reverse of reaction A takes place, xanthone and fluorenone, respectively, being formed, together with isopropyl alcohol.



(B)

It is a true photoreaction, the pinacols being stable to acetone in the dark. Reaction B is an example of the breaking of the —C—C—linkage during a photochemical process in sunlight. There are several ways in which the mechanism of reaction B may be explained; e.g., xanthopinacol may split into free radicals (I) which are dehydrogenated by acetone.

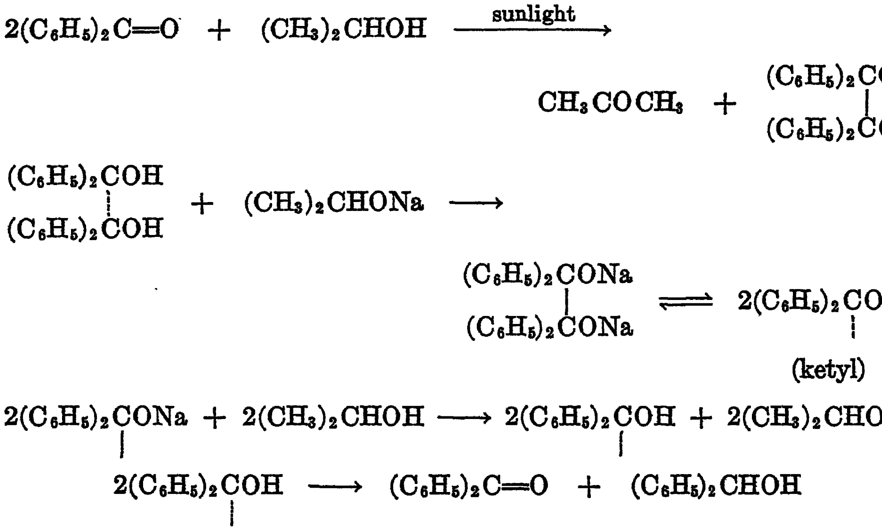


I

It seems not impossible that the reaction between aromatic ketones and isopropyl alcohol (*cf.* scheme A) as well as the reaction between aromatic pinacols and acetone (*cf.* scheme B) leads to an equilibrium which favors, in one case, the formation of pinacols, and in other cases, the formation of aromatic ketones.

B. Reactions of diaryl ketones with alcohol in sunlight in the presence of small amounts of sodium alcoholate

This reaction leads to the formation of benzohydrols, which are formed by the action of sodium alcoholate on benzopinacol, and proceeds, according to Bachmann (3), as follows:



$$2(C_6H_5)_2C{=}O + (CH_3)_2CHOH \xrightarrow{\text{sunlight}} CH_3COCH_3 + \begin{array}{c}(C_6H_5)_2COH\\ |\\ (C_6H_5)_2COH\end{array}$$

$$\begin{array}{c}(C_6H_5)_2COH\\ |\\ (C_6H_5)_2COH\end{array} + (CH_3)_2CHONa \longrightarrow \begin{array}{c}(C_6H_5)_2CONa\\ |\\ (C_6H_5)_2CONa\end{array} \rightleftharpoons \underset{\text{(ketyl)}}{2(C_6H_5)_2\underset{|}{C}ONa}$$

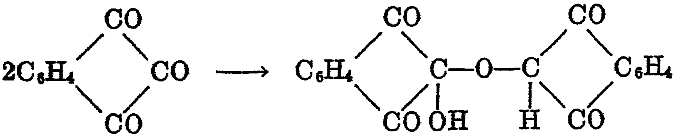
$$2(C_6H_5)_2\underset{|}{C}ONa + 2(CH_3)_2CHOH \longrightarrow 2(C_6H_5)_2\underset{|}{C}OH + 2(CH_3)_2CHONa$$

$$2(C_6H_5)_2\underset{|}{C}OH \longrightarrow (C_6H_5)_2C{=}O + (C_6H_5)_2CHOH \qquad \text{(C)}$$

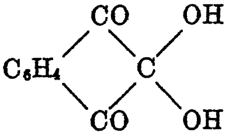
4-Phenylbenzophenone was reduced nearly quantitatively to 4-phenylbenzohydrol. 4-Methylbenzophenone, 4,4′-dimethylbenzophenone, 4-methoxybenzophenone, 4-chloro-4′-methoxybenzophenone, and 4-chlorobenzophenone react also in accordance with scheme C. In the case of benzophenone, the above reaction is suitable for the preparation of benzopinacol (4).

C. Photochemical reduction of 1,2,3-triketones by isopropyl alcohol in the presence of sunlight

The red triketohydrindene (II) is the only substance of this type which has been investigated as far as we know. The reaction does not lead to pinacol formation, but to an ether: namely, the colorless hydrindantin (III) (39) which has previously been obtained in the dark by the action of hydrogen sulfide on ninhydrin (IV). It should be noted that the structure (III) for the reduction product is not beyond doubt and that the formula of the corresponding pinacol is not quite excluded.



II III



IV

Ninhydrin

D. Further methods for the photochemical formation of pinacols

In a number of cases, pinacols have frequently been obtained from diaryl ketones, as by-products, when the ketones were treated with methane derivatives; e.g., when benzophenone is treated with phenylacetic acid (14).

It should also be mentioned that benzopinacol (42) and substituted benzopinacols (43)—e.g., 4,4′,4″,4‴-tetramethylbenzopinacol—may be obtained in excellent yield when the corresponding benzohydrols are exposed to sunlight in the presence of acetone, the reaction proceeding rapidly. Acetone may be replaced by ethyl methyl ketone (42).

$$2C_6H_5CHOHC_6H_5 + CH_3COCH_3 \xrightarrow{\text{sunlight}} (C_6H_5)_2COHCOH(C_6H_5)_2 + CH_3CHOHCH_3$$

III. PHOTOCHEMICAL DEHYDROGENATION EFFECTED BY QUINONES RESULTING IN (*a*) FORMATION OF ALDEHYDES FROM ALCOHOLS, (*b*) FORMATION OF DIARYL KETONES FROM AROMATIC PINACOLS, AND (*c*) FORMATION OF ETHANES FROM THE CORRESPONDING METHANES

Until recently, only the action of *p*-quinones on primary and secondary alcohols had been investigated.

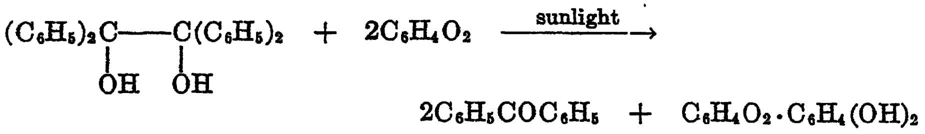
(*a*) *p*-Quinones have the tendency to add hydrogen, changing to hydroquinone; methyl alcohol is thus transformed into formaldehyde (18).

$$C_6H_4O_2 + CH_3OH \xrightarrow{\text{sunlight}} C_6H_4(OH)_2 + HCHO$$

Ethyl alcohol is changed in sunlight into acetaldehyde, and isopropyl alcohol into acetone (9, 11). Dorcas and Forbes (15) selected the action of *p*-benzoquinone on 50 per cent alcohol for actinometric studies. The quantum yield of the reaction has been investigated by Leighton and Forbes (22).

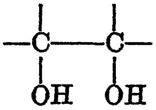
Thymoquinone reacts with ethyl alcohol, acetaldehyde and thymohydroquinone (9, 11) being formed.

(*b*) Recently, photochemical dehydrogenation has been investigated in various fields (43). Whereas pinacol itself is stable to *p*-benzoquinone in sunlight, acetophenonepinacol (V), benzopinacol (VI), 9,10-dihydroxy-9,10-diphenyldihydrophenanthrene (X), and 1,8-diphenylacenaphtheneglycol (XI) are readily dehydrogenated by the action of *p*-benzoquinone in sunlight.



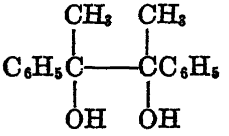
$$(C_6H_5)_2C(OH){-}C(OH)(C_6H_5)_2 + 2C_6H_4O_2 \xrightarrow{\text{sunlight}} 2C_6H_5COC_6H_5 + C_6H_4O_2 \cdot C_6H_4(OH)_2$$

Several derivatives of benzopinacol have also been investigated, e.g., those in which the phenyl group is replaced by VII, VIII, or IX. In all these cases a rupture of a



$$-C(OH){-}C(OH)-$$

linkage is effected and quinhydrone is formed.



V

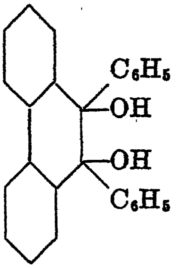
CR_2OHCR_2OH

VI: R = C_6H_5

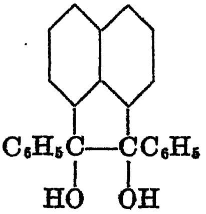
VII: R = p-$CH_3OC_6H_4$

VIII: R = p-$CH_3C_6H_4$

IX: R = p-ClC_6H_4

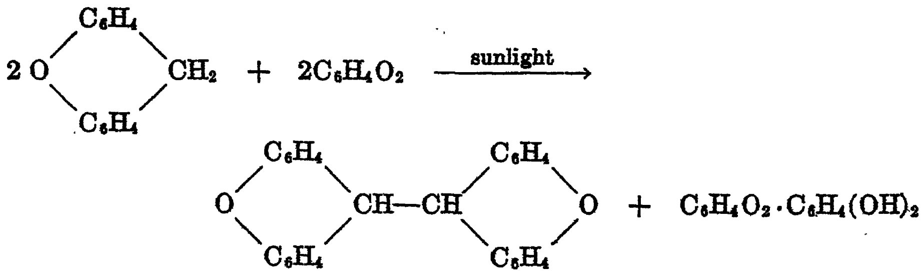


X



XI

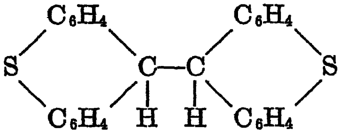
(*c*) The formation of tetraarylethanes from diarylmethanes by the action of a quinone in sunlight has been established by Schönberg and Mustafa (43) in the case of diphenylmethane, fluorene, xanthene, anthrone, and dinaphthopyran; e.g., xanthene leads to the formation of dixanthyl (XIIa).



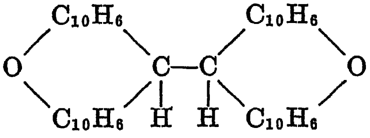
XIIa

Dixanthyl

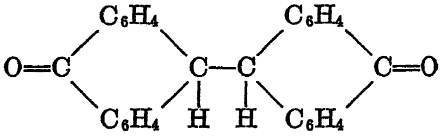
Further, dithiodixanthyl (XIIb) has been produced by the action of *p*-benzoquinone on thioxanthene, and it was also found that in the presence of sunlight (*a*) *s*-tetraphenylethane is produced from diphenylmethane and anthraquinone and also from diphenylmethane and xanthone, (*b*) dianthrone (XIV) is formed from anthrone and xanthone, and (*c*) bisdinaphthoxanthene (XIII) is obtained from dinaphthopyran and xanthone.



XIIb
Dithiodixanthyl



XIIII
Bisdinapthoxanthene

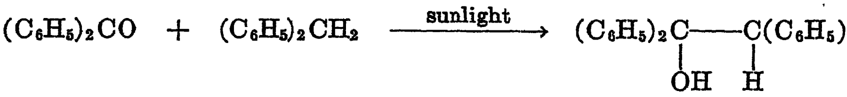


XIV
Dianthrone

Phenanthrenequinone also is able to effect the photochemical dehydrogenation of thioxanthene to dithiodixanthyl (XIIb) (46).

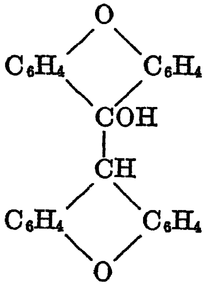
IV. ADDITION REACTIONS BETWEEN KETONES AND METHANES; EXISTENCE OF A PHOTOCHEMICAL EQUILIBRIUM

Paternó and Chieffi (27, 28) have found that diphenylmethane reacts with benzophenone in sunlight with the formation of *s*-tetraphenylethyl alcohol; a similar reaction takes place between benzophenone and dianisylmethane (7).

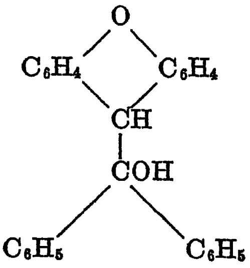


$$(C_6H_5)_2CO + (C_6H_5)_2CH_2 \xrightarrow{\text{sunlight}} \underset{\displaystyle OH}{(C_6H_5)_2C}—\underset{\displaystyle H}{C}(C_6H_5)_2 \quad (D)$$

The reaction between xanthene and xanthone and between xanthene and benzophenone (43) proceeds similarly; the resulting products are XV and XVI. The speed of this reaction is remarkable. 9-Hydroxydixanthyl (XV) separates in crystals of great purity when solutions of xanthene and xanthone in benzene are illuminated by sunlight for a few hours.

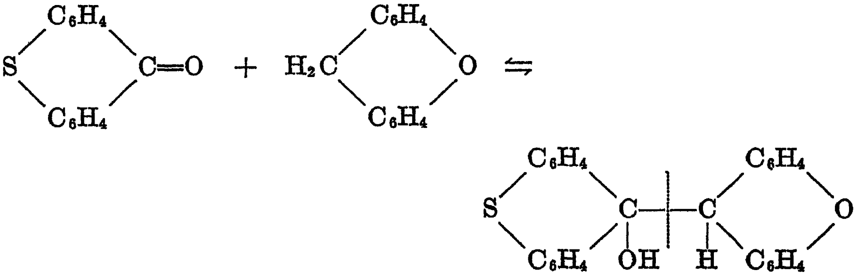


XV
9-Hydroxydixanthyl



XVI

In all these cases, no indication was found that the photoreaction between diaryl ketones and diarylmethanes is reversible, but when the yellow thioxanthone was allowed to react with xanthene (45), both dissolved in benzene, the reaction product (XVII) was formed in small yields and attempts to increase the yield by prolonged illumination were unsuccessful. This failure was due to a photochemical equilibrium.



(E)

XVII

9-Xanthylthioxanthydrol

When 9-xanthylthioxanthydrol (XVII) was suspended in benzene and exposed to sunlight, it partly dissolved and xanthene and thioxanthone were isolated from the solution; moreover, XVII was stable in the dark.

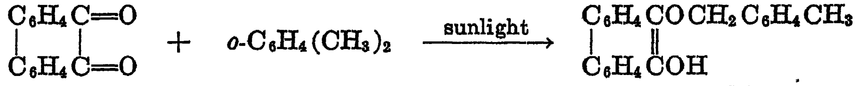
It is believed that the first step in the photodissociation of the carbinol (XVII) is the formation of two free radicals (see broken line in scheme E), which are stabilized by disproportionation. The fact that XVII is colorless, whereas thioxanthone is yellow, is in agreement with the constitution shown above. At 270°C. XVII decomposes into its generators, showing that the ethane linkage in the carbinol is weak.

With regard to the complicated reaction between benzophenone and benzohydrol methyl ether, which may be regarded as a derivative of diphenylmethane, the original literature should be consulted (7).

de Fazi (14) found that when benzophenone was treated with phenylacetic acid in sunlight, an addition product (XVIII) was obtained.

$$(C_6H_5)_2CO + C_6H_5CH_2COOH \xrightarrow{\text{sunlight}} (C_6H_5)_2C(OH)CH(C_6H_5)COOH$$

Benrath and Meyer (6) state that an addition reaction takes place when phenanthrenequinone is exposed to sunlight with *o*- or *p*-xylene (*cf.* XIX). An addition reaction also occurs when phenanthrenequinone is allowed to react with quinaldine.



$$\begin{array}{l}C_6H_4C{=}O\\ |\quad\ \ |\\ C_6H_4C{=}O\end{array} + o\text{-}C_6H_4(CH_3)_2 \xrightarrow{\text{sunlight}} \begin{array}{l}C_6H_4COCH_2C_6H_4CH_3\\ |\quad\ \ \|\\ C_6H_4COH\end{array}$$

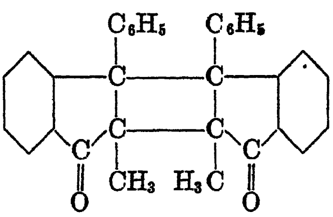
XIX

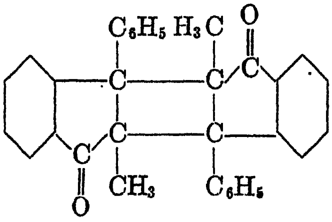
V. PHOTOPOLYMERIZATION REACTIONS OF α,β-UNSATURATED KETONES AND *p*-QUINONES

A large number of these reactions have been described, the majority of them leading to compounds of doubtful constitution.

Chalcone (benzylideneacetophenone), on irradiation in sunlight, forms two dimers: the dimer A, melting at 124°C. (truxinic ketone type), and the dimer B (truxillic ketone type) are believed to be formed (52, 53).

3-Phenyl-2-methylindone gives rise to two different dimeric compounds, believed to be of the following structures (XX and XXI) (13):


XX


XXI

Thymoquinone in ether forms a polymerization product, which on heating dissociates with the formation of thymoquinone (23). 2-Methyl-α-naphthoquinone forms a colorless dimeric product, believed to be a cyclobutane derivative (24).

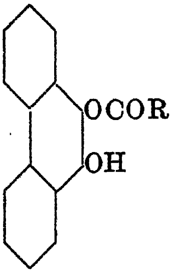
VI. PHOTOCHEMICAL ADDITION OF ALDEHYDES TO (*a*) *p*-BENZOQUINONE, (*b*) PHENANTHRENEQUINONE, (*c*) PHENANTHRENEQUINONIMINE

(*a*) *p-Benzoquinone:* Klinger and Kolvenbach (20) have shown that *p*-benzoquinone reacts with acetaldehyde, yielding $CH_3COC_6H_3(OH)_2$; with isovaleraldehyde, isovalerohydroquinone is obtained (21). Benzaldehyde gives *p*-dihydroxybenzophenone, $C_6H_5COC_6H_3(OH)_2$ (20, 21).

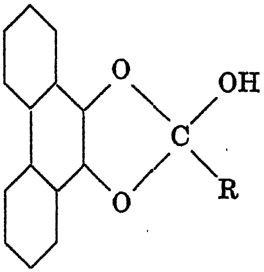
Angeletti and Baledini (2) allowed toluquinone and benzaldehyde to react in benzene in sunlight and obtained, *inter alia,* benzoic acid and toluhydroquinone. The photoreaction between cinnamaldehyde and toluquinone has also been investigated (1).

(*b*) *Phenanthrenequinone:* According to Klinger (19) benzaldehyde reacts with phenanthrenequinone, yielding XXII; acetaldehyde, isovaleraldehyde, anisaldehyde, and cinnamaldehyde react similarly. According to Schönberg and Moubasher (38), the reaction products should be formulated according to XXIII. They are not soluble in an alkaline medium; furthermore, it was found that the photoproducts of phenanthrenequinone and *p*-chlorobenzaldehyde, when treated with diazomethane, yielded a product which must have the constitution XXIV, since on hydrolysis 9,10-dihydroxyphenanthrene and methyl *p*-chlorobenzoate were obtained. Further addition products have been prepared (47), e.g., with 2,4-dimethoxybenzaldehyde, 2-methoxy-1-naphthaldehyde, and cuminaldehyde. Some of the ortho-substituted benzaldehydes—e.g., 2-methoxy-1-naphthaldehyde—were found to react very slowly with phenanthrenequinone when compared with benzaldehyde; this may be due to steric hindrance (ortho effect).

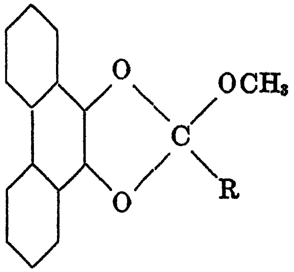
The pyrolysis of the photoproducts obtained by the action of phenanthrenequinone and aldehydes has been investigated. The photoreaction product



XXII
(R = C_6H_5)



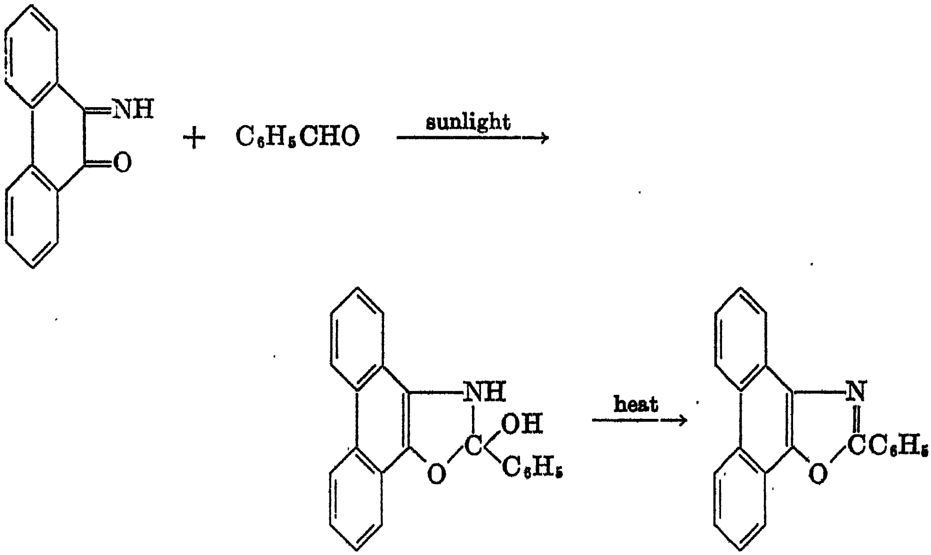
XXIII
(R = C_6H_5)



XXIV
(R = p-ClC_6H_4)

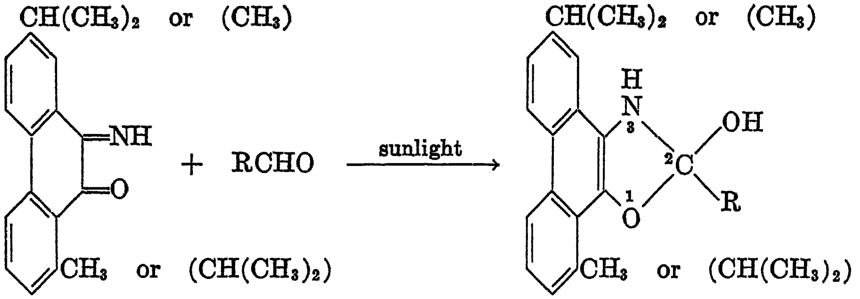
(XXIII) yields benzaldehyde and phenanthrenequinone and a small amount of benzoic acid and fluorenone. For an explanation for the formation of these substances, the original literature should be consulted.

(*c*) *Phenanthrenequinonimine:* The yellow phenanthrenequinonimine (33) reacts with benzaldehyde in sunlight with the formation of the colorless 2-hydroxy-2-phenyl-2,3-dihydrophenanthroxazole (XXV), a substance which on heating yields 2-phenylphenanthroxazole; this result is in agreement with the proposed formula of the photoproduct (XXV). Similar photoreactions were carried out with acetaldehyde, p-anisaldehyde, p-chlorobenzaldehyde, piperonal, and 2-methoxy-1-naphthaldehyde. The reaction proceeds very rapidly with acetaldehyde, benzaldehyde, and p-anisaldehyde.



XXV

Retenequinonimine (XXVI) reacts with acetaldehyde, benzaldehyde, and p-anisaldehyde according to the following scheme (34). The constitution of retenequinonimine has not yet been completely elucidated, as it is not known whether the methyl or the isopropyl group is closer to the carbonyl group.



XXVI

VII. THE ACTION OF SUNLIGHT ON DIPHENYL TRIKETONE

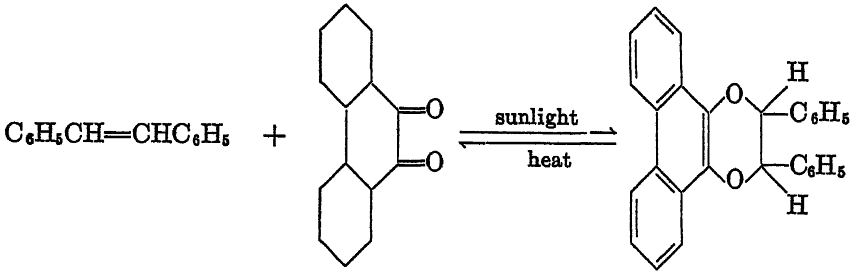
Lately, two methods have been described for converting diphenyl triketone into benzil by dark processes: namely, (*a*) the action of aluminum chloride and (*b*) the action of selenium (precipitated) in the presence of oxygen (40).

It has now been found (47) that when the triketone in benzene solution is exposed to sunlight it is converted into benzil. No reaction takes place in the dark.

$$C_6H_5COCOCOC_6H_5 \xrightarrow{\text{sunlight}} C_6H_5COCOC_6H_5$$

VIII. PHOTOREACTIONS BETWEEN OLEFINS AND 1,2-DIKETONES AND *o*-QUINONES

Schönberg and Mustafa (44, 45, 47) found that when benzil, phenanthrenequinone, or retenequinone (35) in benzene solution was exposed with olefins in sunlight, addition reactions took place, leading to the formation of 1,4-dioxins.



(F)

The velocity of the reaction is remarkable in some cases, e.g., in the case illustrated in scheme F. In general, the velocity with which benzil reacts is much less when compared with phenanthrenequinone. Among the 1,2-diketones which so far have been investigated successfully were only benzil, phenanthrenequinone, and retenequinone. On the other hand a great variety of ethylenes (methyleneanthrone (XLI) and diphenylketone (XLII) among others) were found to be reactive.[1] Further details may be taken from table 1.

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[1] Other ethylenes which were added in sunlight to 1,2-diketones or *o*-quinones are: styrene (XXXII), α,α-diphenylethylene (XXXI), p,p'-dimethoxystilbene (XXXIII), triphenylethylene (XXXIV), α,α-diphenyl-Δ^{α}-propylene (XXXV), α,α-di(*p*-diphenylyl)-β-methylethylene (XXXVI), α-stilbazole (XXXVII), β-chlorostilbene (XXXVIII), 9-benzylidenexanthene (XXXIX), 9-benzylidenethioxanthene (XL), and benzylidenephthalide (XLIII).

TABLE 1

List of photo-addition products obtained from 1,2-diketones and ethylenes

The ethylenes and photoproducts obtained are designated by the same Roman numeral, but the letter "a" is added in the case of the photoproduct; e.g., XXVIIIa is the photoaddition product of XXVIII and benzil

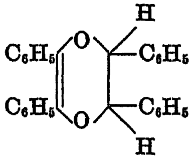
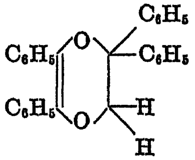
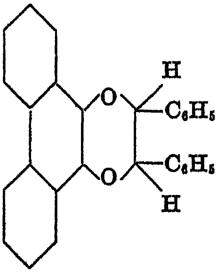
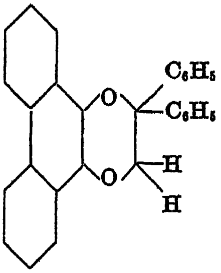
OLEFIN	FORMULAS OF PHOTOADDITION PRODUCTS WITH BENZIL, PHENANTHRENEQUINONE, OR RETENEQUINONE
Reactions with benzil	
$C_6H_5CH{=}CHC_6H_5$ (XXVIII)	 XXVIIIa
$(C_6H_5)_2C{=}CH_2$ (XXIX)	 XXIXa
Reactions with phenanthrenequinone	
$C_6H_5CH{=}CHC_6H_5$ (XXX)	 XXXa
$(C_6H_5)_2C{=}CH_2$ (XXXI)	 XXXIa

TABLE 1—*Continued*

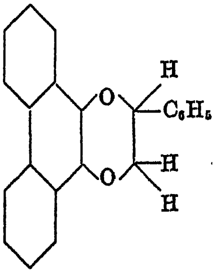
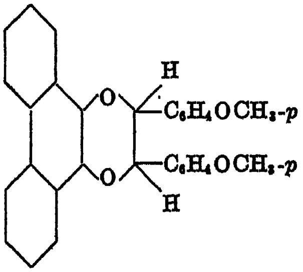
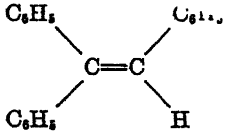
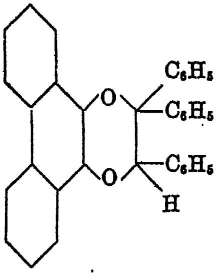
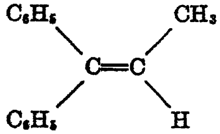
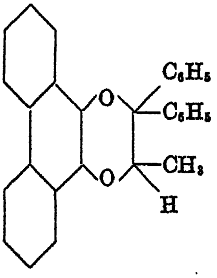
OLEFIN	FORMULAS OF PHOTOADDITION PRODUCTS WITH BENZIL, PHENANTHRENEQUINONE, OR RETENEQUINONE
Reactions with phenanthrenequinone—*Continued*	
$C_6H_5CH{=}CH_2$ (XXXII) p-$CH_3OC_6H_4CH{=}CHC_6H_4OCH_3$-$p$ (XXXIII)	 XXXIIa  XXXIIIa
 $(C_6H_5)_2C{=}CHC_6H_5$ (XXXIV)	 XXXIVa
 $(C_6H_5)_2C{=}CHCH_3$ (XXXV)	

TABLE 1—*Continued*

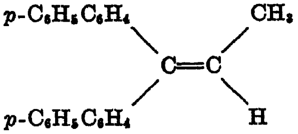
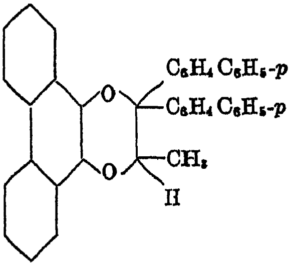
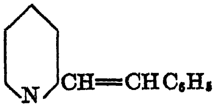
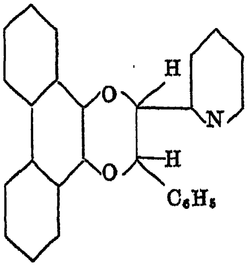
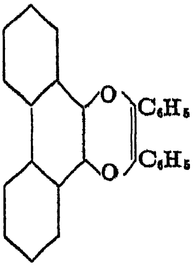
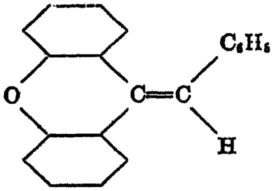
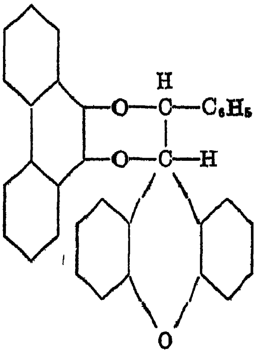
OLEFIN	FORMULAS OF PHOTOADDITION PRODUCTS WITH BENZIL, PHENANTHRENEQUINONE, OR RETENEQUINONE
Reactions with phenanthrenequinone—*Continued*	
 (XXXVI)	 XXXVIa
 (XXXVII)	 XXXVIIa
$C_6H_5CH{=}CClC_6H_5$ (XXXVIII)	 XXXVIIIa*
 (XXXIX)	 XXXIXa

TABLE 1—*Continued*

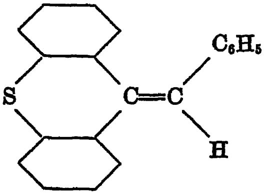
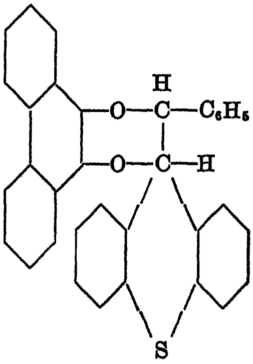
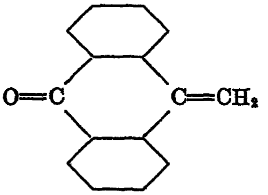
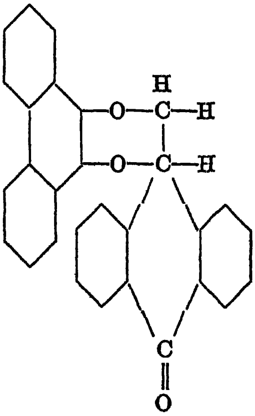
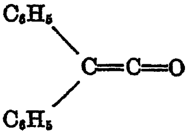
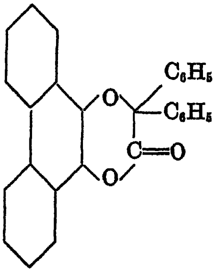
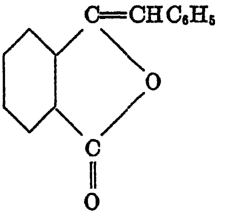
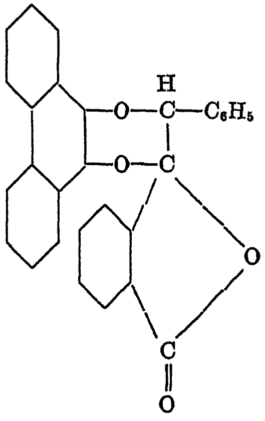
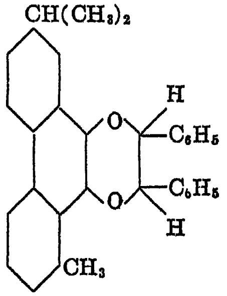
OLEFIN	FORMULAS OF PHOTOADDITION PRODUCTS WITH BENZIL, PHENANTHRENEQUINONE, OR RETENEQUINONE
Reactions with phenanthrenequinone—*Continued*	
 (XL)	 XLa
 (XLI)	 XLIa
 $(C_6H_5)_2C{=}C{=}O$ (XLII)	 XLIIa

TABLE 1—*Concluded*

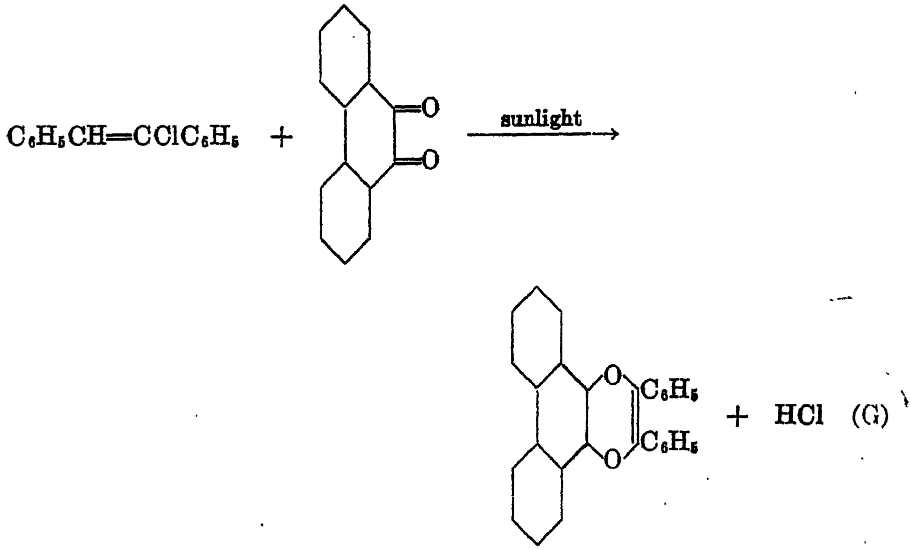
OLEFIN	FORMULAS OF PHOTOADDITION PRODUCTS WITH BENZIL, PHENANTHRENEQUINONE, OR RETENEQUINONE
Reactions with phenanthrenequinone—*Concluded*	
 (XLIII)	 XLIIIa
Reactions with retenequinone	
$C_6H_5CH{=}CHC_6H_5$ (XLIV)	 XLIVa

* Compare scheme G.

In the case of β-chlorostilbene and phenanthrenequinone, on prolonged exposure 2,3-diphenylphenanthro[9,10]-1,4-dioxin (XXXVIIIa) was obtained, hydrogen chloride gas being evolved (*cf.* scheme G on p. 196).

Most of the photoproducts obtained are colorless or light yellow, a fact which is in agreement with their proposed constitution. On heating, the photoaddition products decompose, yielding benzil, phenanthrenequinone, or retenequinone, respectively, and the olefinic product which has been used in their synthesis.

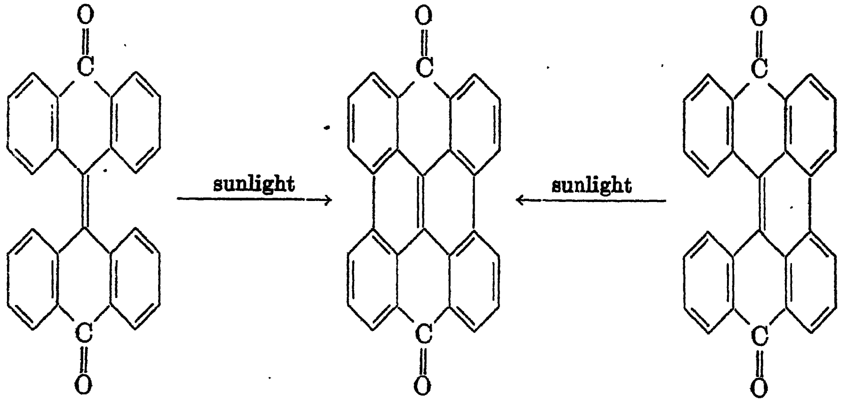
The value of the reactions mentioned in this section lies in the fact that by them a large number of substances have been synthesized which seem to be inaccessible by any other method so far known.



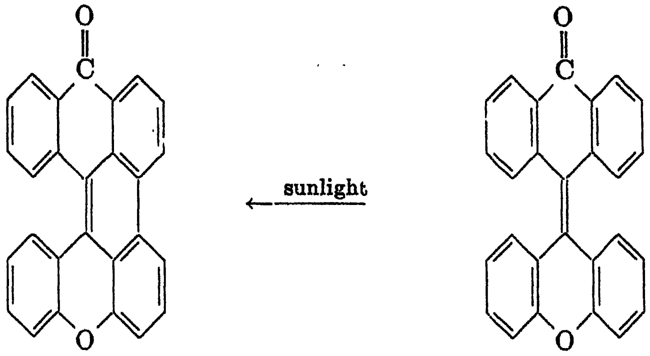
(G)

IX. PHOTOFORMATION OF HIGHLY CONDENSED AROMATIC RING SYSTEMS FROM AROMATIC KETONES BY ELIMINATION OF HYDROGEN ATOMS IN AROMATIC LINKAGES

Examples of these remarkable reactions, which lead to the formation of naphthadianthrone (XLVI) in acetic acid solution, have been observed in the case of XLV by Meyer and coworkers (25); neutral solvents, e.g., xylene or chlorobenzene, may also be used. The same product is also obtained by the irradiation of helianthrone (XLVIa) in acetic acid. Perkin and Yoda (29) have investigated similar cases and have prepared dimethoxynaphthadianthrone. Recently it was shown by Schönberg, Ismail, and Asker (37) that 10-(9′-xanthylene)anthrone (XLVIII) in benzene solution is transformed into oxapenenone (XLVII) in the presence of sunlight.



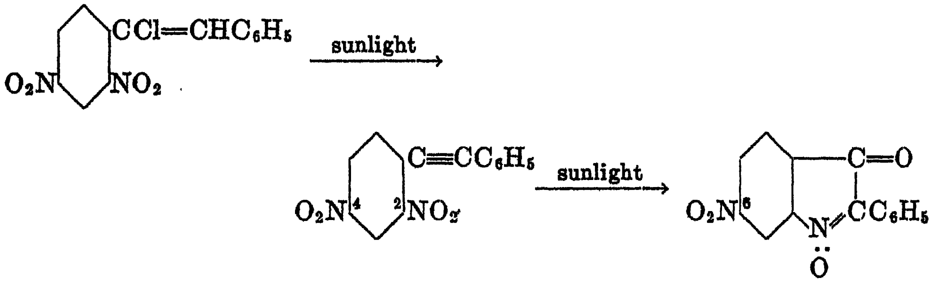
XLV XLVI XLVIa



XLVII XLVIII

X. FORMATION OF HETEROCYCLIC KETONES IN SUNLIGHT

The formation of cyclic ketonic compounds in which the carbon atom of the ketonic group is a member of a five-membered ring system has been observed by Pfeiffer (30). The following reaction was carried out in pyridine and led to the formation of 6-nitro-2-phenylisatogen (XLIX); as an intermediate product 2,4-dinitrotolan is believed to exist. Pfeiffer and Kramer (31) have obtained 6-nitro-2-phenylisatogen (XLIX) from 2,4-dinitrotolan in sunlight.

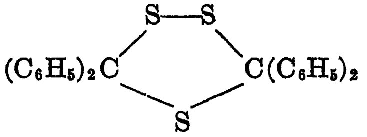


XLIX

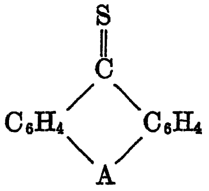
XI. EXPERIMENTS WITH NON-ENOLIZABLE THIOKETONES IN SUNLIGHT

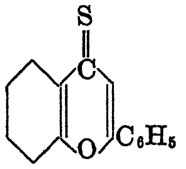
A. Action of oxygen (41)

Thiobenzophenone is extremely sensitive to oxygen under laboratory conditions, being converted into benzophenone, elementary sulfur, sulfur dioxide, and the trisulfide (L) (49, 51). This change takes place even in the dark. On the other hand, 4,4′-dimethoxythiobenzophenone, 4,4′-bisdimethylaminothiobenzophenone, xanthione (LI), and thioxanthione (LII) in benzene solution are distinguished by their stability towards oxygen in the dark, but are converted into the corresponding ketones in sunlight in the presence of oxygen, sulfur and sulfur dioxide being formed. *N*-Phenylthioacridone (LIII), 4-thioflavone (LIV), and 2,6-diphenyldithiopyrone (LV) are stable or almost completely stable in the dark and in sunlight.

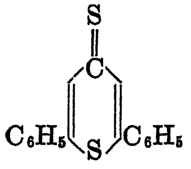


L





LIV



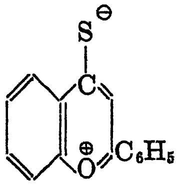
LV

LI: A = O

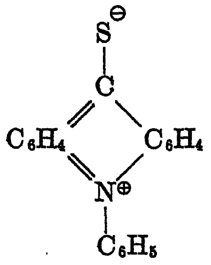
LII: A = S

LIII: A = NC_6H_5

In order to explain the fact that the last three substances are resistant to oxygen in sunlight, it may be pointed out that they are not thioketones proper, but have the structure of zwitter ions (*cf.* LVI and LVII in the case of 4-thioflavone and *N*-phenylthioacridone, respectively) or that the zwitter-ion structure makes a great contribution to their resonance structures.



LVI



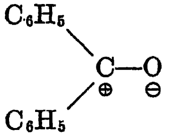
LVII

B. Stability of non-enolizable thioketones towards isopropyl alcohol in the presence of sunlight

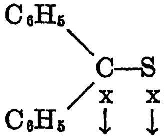
In contrast to the ready reduction of benzophenone and *p,p'*-dimethoxybenzophenone (*cf.* scheme A, page 182), *p,p'*-dimethoxythiobenzophenone, 4,4'-bisdimethylaminothiobenzophenone, xanthione (LI), thioxanthione (LII), *N*-phenylthioacridone (LIII), 2,6-diphenyldithiopyrone (LV), and 4-thioflavone (LIV) were found not to react with isopropyl alcohol after 7 days' exposure (43).

The photostability of thioketones under conditions under which the corresponding ketones react is remarkable, as in the dark thioketones frequently are more active than the corresponding ketones. Thus thiobenzophenone (but not benzophenone) reacts with copper powder (17, 48), diazomethane (8, 36), and phenyl azide (50) in the dark.

It seems possible that the photoactivation of thioketones is different from the photoactivation of ketones; it seems conceivable that, e.g., the photoactivation of benzophenone leads to molecules in which the distribution of electrons is nearly like those in LVIII; on the other hand, the distribution of electrons in the photoactivated thiobenzophenone may be in the direction of LIX.



LVIII



LIX

XII. EXPERIMENTAL NOTES

The authors have found that the most convenient method of carrying out photoreactions in sunlight is to dissolve the components in benzene and to use a Schlenk tube as a reaction vessel (figure 1). This tube is sealed while a stream of dry carbon dioxide is allowed to pass through it. A similar method has been used by Schlenk for the preparation of ketyls in a nitrogen atmosphere (32).

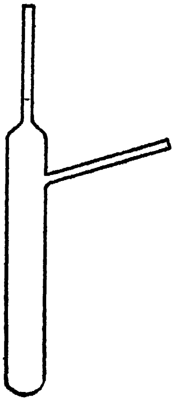


FIG. 1. Schlenk tube used for photoreactions

Special care should be taken that the benzene used is not only thiophene-free, but also free of toluene, which can act as a strong reducing agent, e.g., phenanthrenequinone is transformed to 9,10-dihydroxyphenanthrene by the action of toluene in sunlight (6).

A number of earlier investigators have not stated whether their photoexperiments were carried out in the presence or absence of oxygen. It is necessary to avoid the presence of oxygen in view of the surprising reactivity of oxygen to organic compounds in sunlight (compare the transformation of 2-methylanthraquinone to the corresponding anthraquinonecarboxylic acid in the presence of air (16)).

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THE GAS-PHASE PHOTOCHEMICAL DECOMPOSITION OF THE SIMPLE ALIPHATIC KETONES[1,2]

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I. INTRODUCTION

The studies of the products of the photochemical decomposition of the simple aliphatic ketones such as acetone, ethyl methyl ketone, diethyl ketone, *n*-butyl methyl ketone, and the *n*- and isopropyl ketones have shed much light upon the stabilities and reactions of a number of organic free radicals as well as upon the type of bond that is broken during the photolysis. Acetone, ethyl methyl ketone, and diethyl ketone have been found to dissociate into CH_3, CH_3CO, C_2H_5, and C_2H_5CO radicals. Of these, only C_2H_5CO appears to be very unstable at room temperature. On the basis of the mechanisms proposed to explain the products produced from the photolysis of these three ketones, *a priori* one might have expected *n*-butyl methyl ketone to dissociate into acetyl and butyl radicals to yield final products such as ethane, pentane, octane, etc. However, the data now available indicate that the most important products, acetone and propylene, are formed in a one-step rearrangement reaction. Thus, in the notation of Norrish (41, 52), there are at least two types of primary process in the photolysis of the aliphatic ketones, the size of the groups R and R′ determining the nature of the decomposition:

Type I:

$$RCOR' \rightarrow R'CO + R;\ RCO + R';\ R + R' + CO$$

Type II:

$$RCOCH_2CH_2CH_2R' \rightarrow RCOCH_3 + R'CH{=}CH_2$$

In this article only acetone, ethyl methyl ketone, diethyl ketone, and *n*-butyl methyl ketone are discussed. The first three, decomposing to radicals to a

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large extent, are typical of the Type I reaction, while *n*-butyl methyl ketone is the simplest aliphatic ketone that undergoes the Type II decomposition to a large extent.

II. ACETONE

The photodecomposition of acetone has been the subject, in part or in entirety, of more than fifty original publications (see bibliography), exclusive of many more on the absorption spectrum and fluorescence of acetone. The green fluorescence attributed to acetone prior to 1938 has since been rather conclusively shown to be due to biacetyl (4, 34, 45, 56). The weak blue fluorescence observed when acetone is irradiated with ultraviolet light in the presence of oxygen may be due to acetone, but this has not yet been proved conclusively.

Acetone, as well as other carbonyl compounds, such as other ketones and aldehydes, absorbs light in the ultraviolet, the upper wave-length limit being about 3200–3300 A., or possibly a little higher (62). This upper limit has been shown to consist of a region of discrete absorption (15, 55), probably overlapping a continuum which extends to about 2100 A. (62). A second absorption region beginning at about 1960 A. and extending toward shorter wave lengths has been found (55) to consist of several groups of bands. This region extends to below 1800 A. Acetone has also been shown (20) to exhibit discrete absorption down to about 1300 A. and continuous absorption from 1300 A. at least to 850 A.

Whether there is a real difference between the region of continuous absorption between 2200 and 3050 A. and the so-called region of discrete absorption extending from 3050 A. to about 3300 A. is questionable. Acetone, with twenty-four possible modes of vibration, would be expected to show a complex structure, the extreme types of structure being well-defined vibrational bands and vibrational bands overlapping in such a way as to have the appearance of a continuum. From 2200 to 3050 A. acetone may really have a highly overlapping band structure rather than a truly continuous absorption; from 3050 to 3300 A. the bands that are observed may be the only ones separated enough to permit their being called "bands" rather than part of a "continuous" absorption region.

Studies on the photochemical decomposition of acetone have been made in several of the above-mentioned absorption regions, specifically in the banded region at about 3100 A., in the region of continuous absorption between about 2900 A. and 2200 A., or using approximately monochromatic 2537 A. radiation from mercury arcs. Some work has also been performed using light of about 1900 A. in the photolysis of acetone (37).

The products of the photolysis of acetone have been quite well established. At room temperatures (about 20–30°C.), carbon monoxide, ethane, and biacetyl have been identified by many authors, while methane is also formed to a small extent (71). As the temperature increases the amount of biacetyl formed decreases, until at 60–75°C. it becomes rather unimportant. Further, as the temperature increases above 80–100°C., methane and presumably ethyl methyl ketone are formed in appreciable quantities.

The reaction that takes place when acetone absorbs light, the primary process, probably follows one or more of the courses indicated by equations 1 to 4.

$$CH_3COCH_3 + h\nu \rightarrow CH_3CO + CH_3 \tag{1}$$

$$CH_3COCH_3 + h\nu \rightarrow CH_3COCH_3^* \tag{2}$$

$$CH_3COCH_3 + h\nu \rightarrow C_2H_6 + CO \tag{3}$$

$$CH_3COCH_3 + h\nu \rightarrow CO + 2CH_3 \tag{4}$$

Kirkbride and Norrish (41) first suggested reactions 1, 3, and 4. Spence and Wild (71) considered reaction 2 necessary to explain their data from experiments in the "banded" absorption region around 3130 A. Barak and Style (8) found biacetyl after the photolysis of acetone, proving the necessity of reaction 1. Spence and Wild (70) indicated the importance of reaction 1 with respect to reaction 4 under certain experimental conditions.

Pearson *et al.* (60, 61), using the Paneth mirror technique (57, 58, 59), found that when acetone is photolyzed in a flow system, usually with a mercury arc as a light source, there are formed radicals or atoms which will remove lead, tellurium, and antimony mirrors. These authors analyzed the products, found that metal methyls had been formed, and concluded that methyl radicals were produced in the primary process. Prileshajeva and Terinin (63) found that acetone, when irradiated with magnesium or iron sparks, gave products which removed metal mirrors.

Other evidence that radicals are produced by the photolysis of acetone has been provided by Taylor and Jungers (73) and by Danby and Hinshelwood (17), who found the polymerization of ethylene and acetylene to be induced by acetone when illuminated by ultraviolet light; by Taylor and Rosenblum (74), who found the products of the photolysis of acetone to react with hydrogen; by Gorin (28, 29, 30), who found methyl iodide and acetyl iodide to be formed when acetone is photolyzed in the presence of iodine; by Benson and Forbes (10), who found organic iodides, presumably methyl iodide and acetyl iodide, to be formed when acetone is photolyzed in the presence of iodine; by Feldman, Ricci, and Burton (25), who found lead mirrors to be removed by the photolysis products; and by Feldman, Burton, Ricci, and Davis (24), who determined the rate at which lead mirrors are removed by the products of the photolysis of acetone. Further, experiments on the effect of intensity of absorbed light and the pressures of foreign gases carried out by Leermakers (42), Howe and Noyes (37), Spence and Wild (71, 72), Herr and Noyes (36), Howland and Noyes (38, 39), Saunders and Taylor (69), and Iredale and Lyons (40) all indicate that the production of free radicals during the photolysis of acetone must be quite important.

Of course, one of the stronger pieces of evidence for the free-radical mechanism is derived from the analogy between acetone and ethyl methyl ketone. Norrish (51) first showed that from the latter, nearly equal amounts of the three hydrocarbons ethane, propane, and butane were formed. On the basis of a mechanism such as is suggested in equation 3, only propane would be expected from ethyl methyl ketone. Thus, the data rule out reaction 3 for this ketone and, by analogy, for acetone.

There are only a few ways in which the methyl radical could be formed. Only reaction 1 or 4 would explain the appearance of CH_3 in the primary process. Reaction 2, if this were the primary process, would have to be followed by

$$CH_3COCH_3^* \rightarrow CH_3CO + CH_3 \tag{2a}$$

to explain the appearance of the radicals. Of course, it might be possible to write other, non-radical-producing reactions between 2 and 2a, but this would be rather superfluous. Certainly a reaction like 2a following reaction 2 would also be necessary to explain the presence of biacetyl (8, 70, 71), found to be formed at wave lengths from 2500–3100 A. This product could hardly be expected to be formed by a reaction such as

$$2CH_3COCH_3 \rightarrow C_2H_6 + (CH_3CO)_2$$

instead of a combination of acetyl radicals

$$2CH_3CO \rightarrow (CH_3CO)_2$$

either homogeneously or heterogeneously.

The presence of free radicals in the photolysis of acetone and other information to be presented later makes reaction 3 appear to be a very unlikely reaction, and, of course, it could not explain the presence of biacetyl. However, reactions 1 and 4 can be used to explain the presence of CH_3 and CH_3CO, and the final products ethane, carbon monoxide, and biacetyl.

The reactions of the acetyl radicals are quite numerous. Glazebrook and Pearson (27) and Feldman, Ricci, and Burton (25) have found that after removal of methyl radicals by metal mirrors, if a second mirror farther from the reaction cell is heated, say to 45°C., more metal methyl is formed, presumably by the thermal reactions

$$CH_3CO \rightarrow CH_3 + CO \tag{6}$$

$$CH_3CO + M \rightarrow CH_3 + CO + M \tag{7}$$

The over-all energy of activation for the thermal dissociation of the acetyl radical has been calculated by various authors (9, 30, 33, 36, 50, 65) to be between 10 and 18 kg.-cal. The energy per einstein of 3130 Å. light is about 91 kg.-cal., of 2537 A. light about 113 kg.-cal. If we use the estimate of Rice and Herzfeld (65) of 70 kg.-cal. as the energy necessary to break one C—C bond in acetone, the excess energy retained by the radicals after dissociation is 21 and 43 kg.-cal., respectively, for 3130 and 2537 A. light.

The distribution of this excess energy probably determines, in part, the fate of the acetyl radicals. If no C—H stretching or bending vibrations are excited, the only energy retained by the methyl radical occurs as translational energy, of which there are three degrees. In the acetyl radical, again omitting C—H vibrations, there will be three modes of vibration of the C—C═O skeleton and three degrees of translation. The vibrations of the acetyl radical contain between 7.5 and 15 kg.-cal. (C═O 1700 $cm.^{-1}$, 4.9 kg.-cal.; C—C 900 $cm.^{-1}$, 2.6 kg.-cal.) The remaining energy, appearing as energy of translation, will, from considerations of momentum conservation, be divided roughly inversely as the radical weights. Then the ratio

$$\frac{E_{\text{trans}}(\text{CH}_3)}{E_{\text{trans}}(\text{CH}_3\text{CO})} \simeq \frac{2}{3} \text{ to } \frac{3}{4}$$

To a rough degree of approximation, then, after the dissociation of acetone into acetyl and methyl radicals, the acetyl radical will possess about 10 kg.-cal. as vibrational energy and 3–4 kg.-cal. as translational energy after irradiation with 3130 A. light, a total of 13–14 kg.-cal. As mentioned above, the over-all energy of activation for the thermal decomposition of the acetyl radical has been calculated as 10–18 kg.-cal. The excess energy retained by the acetyl radical, even from acetone photolysis by 3130 A. light, is very close to the energy of activation of the thermal decomposition reaction. We would therefore expect the occurrence, to a small degree after 3130 A. irradiation, and to greater extents after 2537 A. or 1900 A. irradiation, of the spontaneous reaction

$$\text{CH}_3\text{CO}^* \rightarrow \text{CH}_3 + \text{CO} \qquad \text{rate} = \alpha\phi_1 I_a \quad (5)$$

where ϕ_1 is the quantum yield of reaction 1, and α is a proportionality factor which is a function of wave length and temperature. CH_3CO^* is used to designate an activated acetyl radical.

Benson and Forbes (10) have calculated the values of α from their work using 2537 A. light, from the work of Herr and Noyes at 2537 and 3130 A. (36), and from the work of Howe and Noyes at 1900 A. (37).

λ	α
3130	0.07
2537	0.22
1900	0.50

If reaction 5 occurs in such a short time that the time lag has no effect upon the extent to which the reaction takes place, then the sequence of reactions 1 and 5 is equivalent to the reaction

$$\text{CH}_3\text{COCH}_3 + h\nu \rightarrow 2\text{CH}_3 + \text{CO} \qquad \alpha\phi_1 I_a \quad (4)$$

Since there appears to be no evidence to indicate anything but a very short life for CH_3CO^*, there is no difference between the use of a single primary process (reaction 1), followed to some extent by reaction 5, and the use of two primary processes, reactions 1 and 4.

The low value of α (0.07) at 3130 A. probably explains why Gorin (30) found no carbon monoxide when he irradiated acetone with this wave length in the presence of iodine.

Acetyl radicals that have been deactivated by collision, or that are initially formed with insufficient energy to decompose spontaneously, can decompose thermally, since this reaction probably has an upper limit of activation energy of 18 kg.-cal. (see above).

$$CH_3CO + M \rightarrow CH_3 + CO + M \qquad E_{act} = 10\text{–}18 \text{ kg.-cal.} \quad (7)$$
$$\text{rate} = k_7(\text{Ac})(\text{M})$$

Equations 6 and 7 are simplifications of the over-all reactions:

$$CH_3CO + M \rightarrow CH_3CO^* + M \qquad k_{7a}(\text{Ac})(\text{M}) \quad (7a)$$
$$CH_3CO^* \rightarrow CH_3 + CO \qquad k_{7b}(\text{Ac}^*) \quad (7b)$$
$$CH_3CO^* + M \rightarrow CH_3CO + M \qquad k_{7c}(\text{Ac}^*)(\text{M}) \quad (7c)$$

Assuming steady-state conditions for the radicals of these reactions, we arrive at

$$k_{7a}(\text{Ac})(\text{M}) = k_{7b}(\text{Ac}^*) + k_{7c}(\text{Ac}^*)(\text{M})$$

whence

$$(\text{Ac}^*) = \frac{k_{7a}(\text{Ac})(\text{M})}{k_{7b} + k_{7c}(\text{M})}$$

Then the rate of formation of methyl radicals and carbon monoxide by the thermal decomposition of the acetyl radical appears as

$$\frac{d(\text{CO})_7}{dt} = \frac{k_{7b}k_{7a}(\text{Ac})(\text{M})}{k_{7b} + k_{7c}(\text{M})} = \frac{k_{7a}(\text{Ac})(\text{M})}{1 + \frac{k_{7c}}{k_{7b}}(\text{M})}$$

$$k_6 = \frac{k_{7a}(\text{M})}{1 + \frac{k_{7c}}{k_{7b}}(\text{M})}; \qquad k_7 = \frac{k_{7a}}{1 + \frac{k_{7c}}{k_{7b}}(\text{M})}$$

At 25°C. Herr and Noyes (36) found the quantum yield of carbon monoxide formation to increase to a small extent with pressure at 3130 A. ($I_a = 0.5 \times 10^{12}$ quanta/sec./cc.; $P_{\text{acetone}} = 10\text{–}160$ mm.) and at 2537 A. ($I_a = 0.5 \times 10^{12}$ quanta/sec./cc.; $P_{\text{acetone}} = 5\text{–}195$ mm.). At 27°C. Howland and Noyes (39) found a slight increase of Φ_{CO} with pressure at 3130 A. ($I_a = 3 \times 10^{12}$ quanta/sec./cc.; $P_{\text{acetone}} = 54\text{–}203$ mm.) and a slight decrease with pressure at 25–2700 A. ($I_a = 1.2 \times 10^{12}$ quanta/sec./cc., or 12×10^{12} quanta/sec./cc.; $P_{\text{acetone}} = 50\text{–}160$ mm.). At 75°C. Herr and Noyes (36) again found the quantum yield of carbon monoxide formation to increase only slightly with acetone pressures up to 200 mm. at 3130 A. or at 2537 A. At these intensities the constants k_{7a}, k_{7b}, k_{7c} must predict approximate independence of Φ_{CO} with acetone pressure for acetone pressures above 50–100 mm. If (M) in equations 7a and 7c is expressed in millimeters of pressure, $\frac{k_{7c}}{k_{7b}}(\text{M})$ must be considerably greater than unity. In this case

$$\frac{d(\text{CO})_7}{dt} = \frac{k_{7a}k_{7b}}{k_{7c}}(\text{Ac})$$

For this equation to fit the data of Herr and Noyes (36) or of Howland and Noyes (39)

$$\frac{k_{7c}}{k_{7b}}(\mathrm{M}) \geq 10$$

(that is, "very much larger than unity"). At an acetone pressure of 100 mm., $\frac{k_{7c}}{k_{7b}} \geq 0.1$. Assuming the steady-state concentration of acetyl radicals to be nearly independent of pressure from 100–200 mm. acetone, the data of Herr and Noyes indicate $\frac{k_{7c}}{k_{7b}} \simeq 0.3$.

In agreement with experiment we may use equation 6 at moderate and high pressures. At low pressures equation 7 must be used.

$$CH_3CO \rightarrow CH_3 + CO \tag{6}$$

At high pressures of foreign gases, Iredale and Lyons (40) found a considerable increase in the rate of formation of carbon monoxide.

TABLE 1

Effect of carbon dioxide on formation of carbon monoxide

2400–3100 A.; 16 mm. exposures; I_a constant

(Iredale and Lyons (40))

PRESSURE OF ACETONE	PRESSURE OF CO_2	CO FORMED AT N.T.P.
mm.	*mm.*	*cc.*
150	0	0.020
150	0	0.019
150	0	0.020
150	600	0.034
150	620	0.036
150	610	0.034
	Unilluminated	
150	630	0.0005

This increase in the rate of carbon monoxide formation might be attributed either to a decreased quantum yield for acetone decomposed or to the increased decomposition of the acetyl radical by the reaction:

$$CH_3CO + M \rightarrow CH_3 + CO + M \tag{7}$$

The latter possibility suggests that at high foreign gas pressures, biacetyl should cease to be formed, or $C_2H_6/CO \rightarrow 1$. This has been confirmed by Iredale and Lyons, as will be mentioned later.

The data of Iredale and Lyons (40) and those of Howland and Noyes (39) agree quite well with respect to the increased yield of carbon monoxide with the addition of carbon dioxide. The large increase observed is to be attributed to the inefficiency of carbon dioxide as the second body in the deactivating step (equation 7c) as compared to acetone.

The efficiency of the primary process (reaction 1), or processes 1 plus 2, has usually been assumed to be unity. This is not in disagreement with the work of Herr and Noyes (36) on the photolysis of acetone at 3130 A. and at 2537 A.

Unfortunately, experimental difficulties in working at very low pressures are so great that for the photolysis at 25°C. we can say no more than that Φ_{acetone} approaches a value not much less than unity, if less than unity at all (figures 1 and 2).

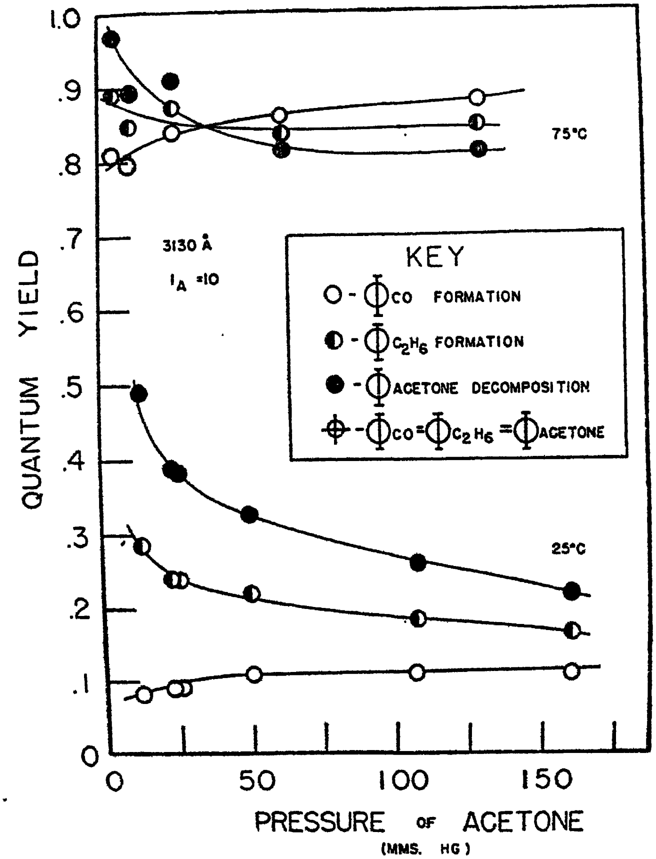

FIG. 1. Effect of pressure and temperature on quantum yields at 3130 A. (Herr and Noyes (36).)

The explanation of the low over-all quantum yield of acetone decomposed is generally given by the recombination reaction (equation 8).

	E_{act}		
$CH_3CO + CH_3 \rightarrow CH_3COCH_3$			(8)
$2CH_3 \rightarrow C_2H_6$	8 kg.-cal.	(ref. 65)	(9′)
$CH_3 + CH_3CO \rightarrow C_2H_6 + CO$	6 kg.-cal.	(ref. 69)	(10)
$2CH_3CO \rightarrow (CH_3CO)_2$	Low		(11′)

To compete, at moderate pressures and at room temperature, with reactions 9′, 10, and 11′, reaction 8 has been assigned an activation energy as low as zero. If reactions 8, 9′, 10, and 11′ were all gas-phase reactions at moderate pressures

(10–200 mm.) and moderate absorbed intensities (10^{11}–10^{13} quanta/cc./sec.) at room temperatures, we might expect them to have about equal probabilities except for steric factors, since all activation energies are quite small. However, at very low pressures reaction 8 appears unimportant, the quantum yield of acetone decomposition approaching a value certainly not much less than unity. At higher pressures, Φ_{acetone} decreases, reaction 8 becoming quite important. We can reword these statements and say that at low pressures reactions 6, 7, 8, and 10 must be relatively unimportant, because of the low values of Φ_{CO}, while

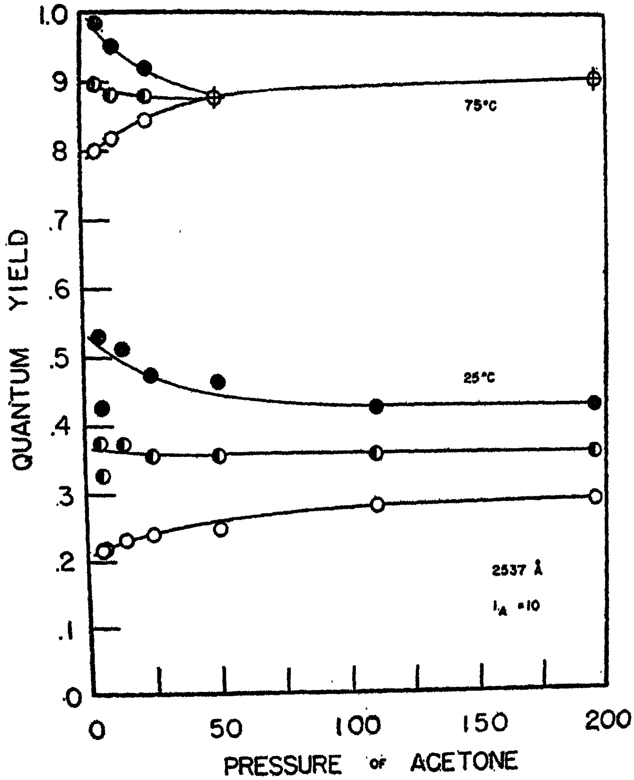

FIG. 2. Effect of pressure and temperature on quantum yields at 2537 A. (Herr and Noyes (36).)

reactions 9′ and 11′, neither being considered to imply the kinetics of the radical combinations, remain important. Reaction 9′ must remain important because of the increase of ethane formation as the pressure decreases. Reaction 11′ must remain important to explain the absence of acetyl radicals for the formation of carbon monoxide.

The obvious conclusion to consider concerning reactions 9′ and 11′ is that these are both wall reactions. This would explain high values of ethane and biacetyl (calculated by a material balance), low values of carbon monoxide, and high values for the ratio C_2H_6/CO at low pressures.

Also, with diffusion becoming unimportant at higher pressures, reactions 8 and 10 become more important, requiring these reactions to be homogeneous. Since the over-all quantum yield of acetone decomposition does not continue to decrease with increasing pressure, but, rather, attains a nearly constant value (figures 1 and 2), reaction 8 cannot be pressure dependent. The same reasoning holds for reaction 10; that is, Φ_{CO} and $\Phi_{C_2H_6}$ both attain approximately steady values as the pressure increases. These considerations assure us of the importance of the following reactions:

	E_{act}	*Rate*	
$CH_3CO + CH_3 \rightarrow CH_3COCH_3$	0 kg.-cal.	$k_8(Ac)(CH_3)$	(8)
$CH_3 \xrightarrow{wall} \frac{1}{2}C_2H_6$	0–8 kg.-cal.	$k_9(CH_3)/P$	(9)
$CH_3CO + CH_3 \rightarrow C_2H_6 + CO$	6 kg.-cal.	$k_{10}(Ac)(CH_3)$	(10)
$CH_3CO \xrightarrow{wall} \frac{1}{2}(CH_3CO)_2$		$k_{11}(Ac)/P$	(11)

There are arguments both for and against reaction 10. In favor of this reaction or of a reaction unlike 6 or 7 is the low energy of activation, 8 kg.-cal., found by Gorin (30) for carbon monoxide formation when the photolysis is performed in the presence of iodine, as contrasted to most estimates of 14–18 kg.-cal. for reactions 6 or 7 (see above). Further, a reaction of this type might explain the low quantum yield of biacetyl formation above 60–75°C., assuming the acetyl radical to be stable enough at this temperature to exist in appreciable concentration. One other argument in favor of reaction 10 is the low yield of carbon monoxide (only about 8 per cent of methyl iodide) and the high yield of acetyl iodide found in the photolysis of acetone in the presence of iodine even at 100°C., at which temperature an activation energy of 14–18 kg.-cal. for reactions 6 and 7 would lead one to expect 60–90 per cent thermal decomposition of the acetyl radical. In contradiction of the necessity of reaction 10 is the geometry, which would require a possibly unexpected arrangement of hydrogen atoms in the acetyl radical to allow for the formation of ethane upon collision with a methyl radical. A high efficiency and/or a low activation energy for the reaction

$$CH_3CO + I_2 \rightarrow CH_3COI + I$$

would explain the low quantum yield of carbon monoxide formation in the presence of iodine. Spence and Wild (71) reported C_2H_6/CO of about 1.08 at 51°C. and 80 mm. pressure. This indicates that biacetyl is formed in small quantities at 50°C. Fluorescence data also indicate biacetyl formation at 85°C. (46). Experiments by Anderson and Rollefson (6) on the photolysis of acetone at about 50 mm. pressure indicate that traces of biacetyl are formed even at 110°C. At this pressure, as can be seen in table 8, biacetyl is quite important at 65°C.

In summary of the value of reaction 10 it can probably be said that this reaction is useful, but there is no irrefutable proof of its necessity. The presence or absence of biacetyl above 60–75°C. can be explained if reaction 10 is omitted, since reaction 6 or 7 predicts 60–90 per cent decomposition of the acetyl radical at about 100°C., thus allowing for the presence of a small concentration of this radical at this temperature. At low pressures biacetyl might be expected at 60–75°C., or at high pressures and very high absorbed intensities. However, at moderate intensities, reactions 6 and 7 will be favored over reactions 11′ or 12, each of which is a function of the square of the acetyl radical concentration.

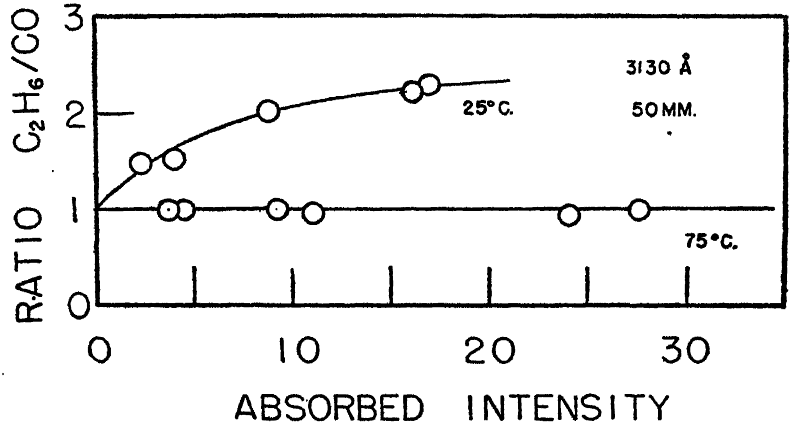

FIG. 3. Effect of absorbed intensity and temperature on C_2H_6/CO ratio at 3130 A. (Herr and Noyes (36).)

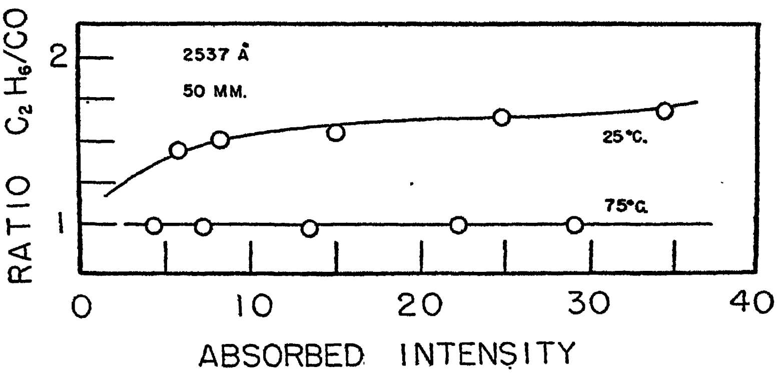

FIG. 4. Effect of absorbed intensity and temperature on C_2H_6/CO ratio at 2537 A. (Herr and Noyes (36).)

Although reactions 1 and 4–11 explain most of the experimental data obtained from the photolysis of acetone at room temperature with either "banded" light around 3100 A. or continuum light from 3000 A. to 2200 A., a few other reactions appear to be necessary in considerations of the effects of absorbed intensity and cell size.

For a given pressure at 25°C. the equations so far considered will not explain the experimentally observed increase in the C_2H_6/CO ratio with absorbed intensity (36). The increase in the biacetyl production, a parallel effect, can only be explained if this product can be formed homogeneously at 3130 A. or in the 2200–3000 A. continuum.

$$2CH_3CO + M \rightarrow (CH_3CO)_2 + M \qquad k_{12}\,(Ac)^2M^y \qquad (12)$$

or

$$2CH_3CO \rightarrow (CH_3CO)_2 \qquad k_{11'}(Ac)^2 \qquad (11')$$

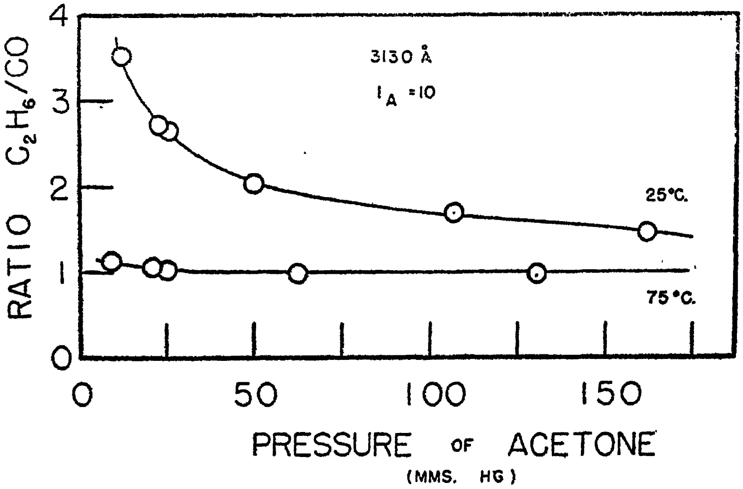

FIG. 5. Effect of pressure and temperature on C_2H_6/CO ratio at 3130 A. (Herr and Noyes (36).)

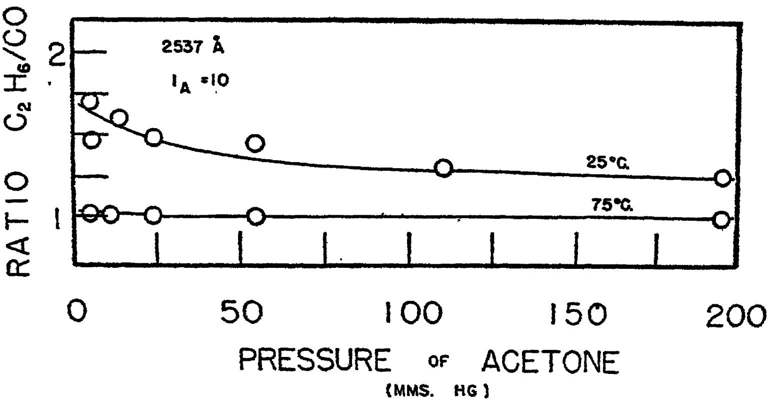

FIG. 6. Effect of pressure and temperature on C_2H_6/CO ratio at 2537 A. (Herr and Noyes (36).)

That either reaction 12 or 11′ is important is shown in figures 5 and 6. At pressures of 150 mm. of acetone the biacetyl yield has not decreased to zero, or the ratio C_2H_6/CO has not decreased to unity. These data are for moderate absorbed intensities.

At higher intensities, the equations so far considered appear to be in agreement with the experimental results. Studies on the reactions of the acetyl radical carried out by Anderson and Rollefson (6) using 2537 A. radiation, with intensities about one hundred times those used by Noyes *et al.*, have shed considerable light on this subject. The reactions of the acetyl radical in the presence of biacetyl are dissociation to methyl and carbon monoxide or recombination. If one of these reactions were homogeneous and the other heterogeneous, the presence of considerable quantities of nitrogen would effect a change in the carbon monoxide yield. That this is not the case is shown in table 2. Hence, the formation of biacetyl at high intensities and moderate pressures and the decomposition of the acetyl radical must both be either homogeneous or heterogeneous. Since reaction 6 or 7 has been shown to occur mainly in the gas phase, the formation of biacetyl at high intensities must be homogeneous also, to a large extent. These considerations, of course, do not rule out heterogeneous reactions at low pressures.

TABLE 2

Reactions of the acetyl radical

Biacetyl; 25°C.; time of run, 20 min.

(Anderson and Rollefson (6))

Initial $(CH_3CO)_2$ (moles × 10^5)	6.88	6.95
Initial N_2 (moles × 10^5)	0	26.42
$\Delta P/P_i$ per cent	32.9	29.8
Products (moles × 10^5):		
CO	2.41	2.37
CH_4	0.07	
C_2H_6	0.69	0.61
CO yield	0.351	0.341

Spence and Wild (71, 72) have shown that much of the above mechanism must be correct. These authors (71) found that the C_2H_6/CO ratio increases with rate of decomposition of acetone at low rates of decomposition, in agreement with equation 12 or 11′. As the intensity is further increased, the ratio apparently passes through a maximum value for banded absorption and then decreases somewhat.

An explanation of this maximum comes from consideration of equations 1, 5, 6, 8, 9′, and 11′:

$$CH_3COCH_3 + h\nu \rightarrow CH_3CO + CH_3 \qquad \phi_1 I_a \qquad (1)$$

$$CH_3CO \rightarrow CH_3 + CO \qquad \alpha\phi_1 I_a \qquad (5)$$

$$CH_3CO \rightarrow CH_3 + CO \qquad k_6(\mathrm{Ac}) \qquad (6)$$

$$CH_3CO + CH_3 \rightarrow CH_3COCH_3 \qquad k_8(\mathrm{Ac})(CH_3) \qquad (8)$$

$$CH_3 + CH_3 \rightarrow C_2H_6 \qquad k_{9'}(CH_3)^2 \qquad (9')$$

$$CH_3CO + CH_3CO \rightarrow (CH_3CO)_2 \qquad k_{11'}(\mathrm{Ac})^2 \qquad (11')$$

At high intensities and high pressures (>100 mm.), diffusion being relatively unimportant, these reactions can be used to explain most of the data.

From steady-state considerations:

$$\phi_1 I_a(1 - \alpha) = k_6(\mathrm{Ac}) + k_8(\mathrm{Ac})(\mathrm{CH_3}) + k_{11'}(\mathrm{Ac})^2$$

$$\phi_1 I_a(1 + \alpha) = -k_6(\mathrm{Ac}) + k_8(\mathrm{Ac})(\mathrm{CH_3}) + k_{9'}(\mathrm{CH_3})^2$$

From this

$$(\mathrm{Ac}) = \frac{-(k_8(\mathrm{CH_3}) + k_6) + \{(k_8(\mathrm{CH_3}) + k_6)^2 + 4k_{11'}(1 - \alpha)\phi_1 I_a\}^{\frac{1}{2}}}{2k_{11'}}$$

$$\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}t} = \frac{k_{9'}(\mathrm{CH_3})^2}{2}$$

$$\frac{\mathrm{d}(\mathrm{CO})}{\mathrm{d}t} = k_6(\mathrm{Ac}) + \alpha\phi_1 I_a$$

$$\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})} = \frac{k_{9'}(\mathrm{CH_3})^2}{2\{k_6(\mathrm{Ac}) + \alpha\phi_1 I_a\}}$$

Now

$$(\mathrm{CH_3}) = \frac{-k_8(\mathrm{Ac}) + [k_8^2(\mathrm{Ac})^2 + 4k_{9'}\{k_6(\mathrm{Ac}) + \phi_1 I_a(1 + \alpha)\}]^{\frac{1}{2}}}{2k_{9'}}$$

Then

$$\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})} = \frac{\{-k_8(\mathrm{Ac}) + [k_8^2(\mathrm{Ac})^2 + 4k_{9'}\{k_6(\mathrm{Ac}) + \phi_1 I_a(1 + \alpha)\}]^{\frac{1}{2}}\}^2}{8k_{9'}\{k_6(\mathrm{Ac}) + \alpha\phi_1 I_a\}}$$

If we use as a first approximation $(\mathrm{Ac}) = KI_a$ at high intensities

$$\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})} = \frac{\{-k_8 K I_a + [k_8^2 K^2 I_a^2 + 4k_{9'}\{k_6 K I_a + \phi_1 I_a(1 + \alpha)\}]^{\frac{1}{2}}\}^2}{8k_{9'}\{k_6 K I_a + \alpha\phi_1 I_a\}}$$

let

$$-k_8 K = C_1; \qquad 4k_{9'}\{k_6 K + \phi_1(1 + \alpha)\} = C_2$$

$$8k_{9'}(k_6 K + \alpha\phi_1) = C_3$$

$$\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})} = \frac{[C_1 I_a + \{C_1^2 I_a^2 + C_2 I_a\}^{\frac{1}{2}}]^2}{C_3 I_a}$$

$$\frac{\mathrm{d}}{\mathrm{d}I}\left(\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})}\right) = \frac{\mathrm{d}}{\mathrm{d}I}\left[\frac{2C_1^2 I_a}{C_3} + \frac{2C_1}{C_3}\{C_1^2 I_a^2 + C_2 I_a\}^{\frac{1}{2}} + C_2\right]$$

$$= \frac{2C_1^2}{C_3} + \frac{C_1}{C_3}\frac{2C_1^2 I_a + C_2}{(C_1^2 I_a^2 + C_2 I_a)^{\frac{1}{2}}}$$

For large values of I_a ; i.e., $I_a \ll I_a^2$

$$\frac{\mathrm{d}}{\mathrm{d}I}\left(\frac{\mathrm{d}(\mathrm{C_2H_6})}{\mathrm{d}(\mathrm{CO})}\right) \simeq \frac{4C_1^2}{C_3} + \frac{C_2}{C_3 I_a}$$

Thus, at very high values of the absorbed intensity, the ratio $d(C_2H_6)/d(CO)$, or $(C_2H_6)/(CO)$, appears as a decreasing function of the intensity (see table 3). Although the assumption that for high intensities $(Ac) = KI_a$ is only approximate, it is probably close enough to allow the above reasoning a considerable validity.

TABLE 3

Influence of intensity on C_2H_6/CO *ratio*

Full mercury arc

(Spence and Wild (71))

ACETONE PRESSURE	TEMPERATURE	TIME	SCREEN TRANSMISSION	ACETONE DECOMPOSED		C_2H_6/CO
				Per cent	Yield per 10 hr. (cc. at N.T.P.)	
mm.	*°C.*	*hr.*	*per cent*			
163	22.0	17	100	30.7	2.81	1.62
163	19.5	2.5	100	4.2	2.60	1.62
164	20.5	3	100	6.0	3.13	1.59
162	21.0	19	60	17.4	1.42	1.76
165	20.5	15	40	9.7	1.01	1.84
161	21.0	24	18	7.7	0.49	1.98
161	20.0	48	7	5.6	0.19	1.87
164	19.0	67	2.5	2.8	0.066	1.56

TABLE 4

Photodecomposition in the region of continuous absorption. Influence of light intensity and acetone pressure

(Spence and Wild (71))

ACETONE PRESSURE	TEMPERATURE	TIME	SCREEN TRANSMISSION	ACETONE DECOMPOSED		C_2H_6/CO	$(p^{1/4})$
				Per cent	Yield per 10 hr. (cc. at N.T.P.)		
mm.	*°C.*	*hr.*	*per cent*				
160	20.0	48	100	43	1.37	2.44	(2.42)
160	19.5	21.5	100	16.9	1.19	2.44	(2.42)
161	21.0	17	100	16.2	1.46	2.47	(2.42)
160	20.0	21	60	9.3	0.68	2.18	
160	21.0	25.5	29	5.2	0.31	2.23	
160	20.0	41	12	3.1	0.115	1.87	
160	21.0	48	12	2.5	0.079	1.63	
70	20.0	24	100	21.6	0.598	1.94	(1.97)
70	20.0	18	100	19.4	0.719	1.92	(1.97)
32	20.0	26	100	31.1	0.367	1.62	(1.62)

Spence and Wilde (71, 72) also found the C_2H_6/CO ratio to increase with pressure at high intensities in the region of continuous absorption (see table 4).

Their data are in agreement with equation 12 or 11′; that is, at high intensities, the homogeneous formation of biacetyl becomes increasingly important as the pressure increases. If reaction 12 is used to explain the data of Spence and

Wild (71), the exponent y must be considerably less than 0.5. In the absence of any other effects, a value of 0.25 for y fits the data quite well.

The effects of high pressures of foreign gases upon the photolysis of acetone at 18–20°C. have been studied by Iredale and Lyons (40) at 2400–3100 A. (see table 5). These studies show that at very high pressures biacetyl ceases to be formed, the ratio C_2H_6/CO becoming unity. Thus, the homogeneous formation of biacetyl becomes of less importance than the decomposition of the acetyl radicals or the re-formation of acetone.

Spence and Wild (71, 72) have proposed a mechanism involving long-lived excited acetone molecules to explain their data for decomposition in the banded absorption region. These authors have considered the following sequence of reactions for absorption of light around 3100 A., and consider them to occur simultaneously with some of the radical reactions.

TABLE 5

Effects of high pressures of foreign gases upon the photolysis of acetone

T = 18–20°C.

(Iredale and Lyons (40))

ADDED GAS	PRESSURE OF ADDED GAS	PRESSURE OF ACETONE	BIACETYL FORMED
	mm.	*mm.*	*mg.*
		150	2.0
		155	2.0
		140	2.0
CO_2	620	150	0
CO_2	590	150	0
N_2	600	150	0
N_2	640	140	0
N_2	640	150	0

$$CH_3COCH_3 + h\nu \rightarrow CH_3COCH_3^* \qquad (2)$$

$$CH_3COCH_3^* \xrightarrow[\text{low pressures}]{10^{-8}\text{ sec.}} CH_3 + CH_3CO \qquad (2a)$$

$$CH_3COCH_3^* + M \rightarrow CH_3COCH_3^{**} + M \qquad (13)$$

$$CH_3COCH_3^{**} \xrightarrow{10^{-2}\text{ sec.}} C_2H_6 + CO \qquad (14)$$

$$CH_3COCH_3^{**} \rightarrow CH_3COCH_3 \qquad (15)$$

$$CH_3COCH_3^{**} \xrightarrow{\text{wall}} C_2H_6 + CO \qquad (16)$$

$$CH_3COCH_3^{**} + CH_3CO \rightarrow CH_3COCH_3 + CH_3 + CO \qquad (17)$$

Reactions 13–15 are the most important at high acetone pressures.

Probably the easiest way to present the foundations of these reactions is to use exactly the arguments of Spence and Wild (72). These authors considered four effects which they believed could not be explained on the basis of the free-radical mechanism for acetone decomposition in the banded region.

(*a*) At constant I_a, the quantum yield and C_2H_6/CO ratio decrease with increase of pressure. This effect is much more pronounced than in the short-wave region and the data suggest that the quantum yield approaches unity as $P \rightarrow 0$.

(*b*) At constant pressure, the quantum yield is independent of I_a, in contrast to its behavior in the "continuum" region. On the other hand, the C_2H_6/CO ratio increases with increase of intensity as in the "continuum" region.

(*c*) Influence of the wall: At moderate pressures (160 mm.), a reduction in the mean molecular distance from the wall by the use of narrower vessels causes C_2H_6/CO to decrease toward unity. This effect would also be expected to become noticeable when the pressure is reduced in any given vessel, but actually, the ratio increases rapidly with fall of pressure.

(*d*) Effect of mixed light: At high intensities, the C_2H_6/CO value obtained with "continuum" light alone is reduced by the presence of light in the banded region to an extent which is considerably greater than would be expected on a purely additive basis (71).

Almy *et al.* (3, 4, 5) have shown that the energy of fluorescence of biacetyl must have been imparted to the biacetyl by energy transfer; that is, biacetyl is excited by collisions of the second kind. Although Almy *et al.* state that their data do not allow them to postulate a unique method of energy transfer, it is possible that the collision of the second kind involves excited acetone molecules and normal biacetyl molecules.

The work of Herr and Noyes (36) indicates that at acetone pressures of 160 mm., biacetyl production is very important when the photolysis is performed using 3130 A. light. The intensity of biacetyl fluorescence increases with time (3) and becomes measurable a few seconds after the start of illumination. Since the partial pressure of biacetyl that can be built up after a few seconds' illumination must be extremely small, if excited acetone molecules are the source of the energy transferred to biacetyl, it follows that the former must be relatively long-lived. Spence and Wild (72) consider this consistent with their data for mixed light (condition *d*) and with the influence of diameter at high pressures.

The diameter effect[3] seems to be open to considerable question. In their earlier work Spence and Wild (71) used a cylindrical cell of about 3.6 cm. inside diameter. As the mean displacement of molecules from the wall these authors used a value of 0.6 cm., i.e., one-third the radius. Assuming that the 7 cm. length of the cell would allow it to be called a long cell, this value would be correct. More accurately $d = 0.52$ cm. The annular cell, 1 cm. between walls,

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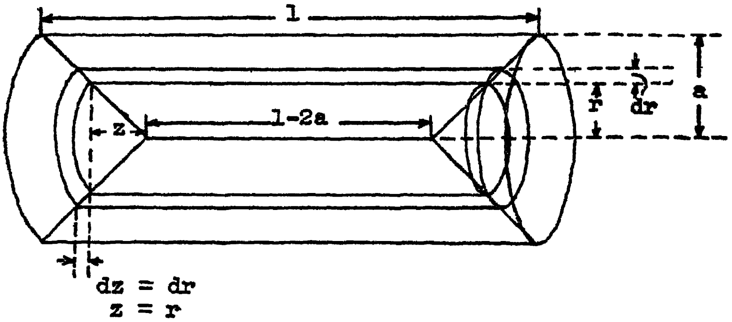
[3] Determination of the mean distance of molecules from the walls of a cylindrical cell with plane, parallel windows.

For simplicity we divide cells into two categories: (*1*) $l/2 \geq a$; (*2*) $l/2 \leq a$. In both cases we consider the distance to the nearest wall and require that all parts of the volume element be the same distance from the nearest wall.

Case 1. $l/2 \geq a$:

c = concentration in arbitrary units, and
d = mean, or average, molecular distance.

(Footnote 3—continued)



$$dN = c\,dV = 2\pi c\{r(l - 2a + 2r) + r^2\}\,dr$$

The distance to the nearest wall is $(a - r)$

$$d = \frac{2\pi c \int_0^a \{r(l - 2a + 2r) + r^2\}(a - r)\,dr}{\pi a^2 lc}$$

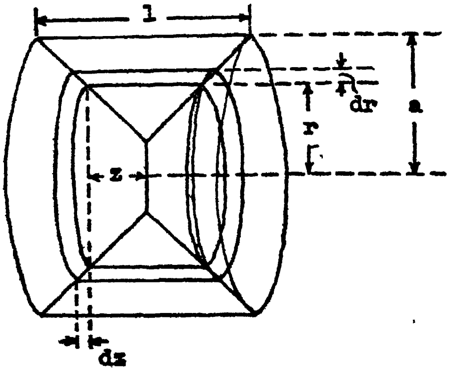
$$= \frac{2\int_0^a \{-3r^3 + r^2(5a - l) + r(la - 2a^2)\}\,dr}{a^2 l}$$

$$= \frac{\left[-\frac{3r^4}{2} + \frac{2r^3}{3}(5a - l) + r^2(la - 2a^2)\right]_0^a}{a^2 l}$$

$$d = -\frac{a^2}{6l} + \frac{a}{3} \qquad \text{as } l \rightarrow \infty, \qquad d \rightarrow \frac{a}{3}$$

$$\frac{l}{2} = a, \qquad d = \frac{l}{8} = \frac{a}{4}$$

Case 2. $l/2 \leqslant a$:



used in the later work (72) had a mean displacement of 0.25 cm. In using the data of Herr and Noyes (36), Spence and Wild used a value of the mean displacement of 1 cm. (1.2 cm. radius cell) instead of about 0.39. Therefore, the low values of C_2H_6/CO obtained by Spence and Wild (71, 72) with cells of mean displacement of 0.25 and 0.52 cm., are very much lower than those obtained by Herr and Noyes, at the same pressure and intensity, in a cell of intermediate mean displacement of 0.39 cm. The ratios obtained by Herr and Noyes (36) were checked quite closely by Howland and Noyes (39), using a cell the same size as that used by the former authors. Also, using a cell of 2.2 cm. inside diameter and 1 cm. long (mean displacement 0.18 cm.), Howland and Noyes found the diameter effect to be small, but in the opposite direction to that found by Spence and Wild, i.e., Howland and Noyes found that the ratio C_2H_6/CO increased as the mean displacement decreased.

A possible explanation of part or all of the difference in the C_2H_6/CO ratio

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(*Footnote 3—continued*)

Here $r = a - 2 + z$.

$$dN = \pi c\left\{\left(a - \frac{l}{2} + z\right)^2 + 2\left(a - \frac{l}{2} + z\right)z\right\}dz$$

$$d = \frac{2\pi c\int_0^{l/2}\left\{\left(a - \frac{l}{2} + z\right)^2 + 2\left(a - \frac{l}{2} + z\right)z\right\}dz}{\pi c a^2 l}$$

$$d = \frac{l^3 - 8l^2 a + 24a^2 l}{96a^2}$$

$$\text{as } a \longrightarrow \infty \qquad d \longrightarrow l/4$$

$$a = \frac{l}{2} \qquad d = \frac{l}{8} = \frac{a}{4}$$

The cell used by Spence and Wild (71) had dimensions $l = 7$ cm. and $a = 1.8$ cm. This cell is analyzed according to *Case 1*: i.e., $l/2 > a$.

$$d = \frac{-a^2}{6l} + \frac{a}{3} = -0.08 + 0.6 = 0.52 \text{ cm.}$$

The cell used by Herr and Noyes (36) had dimensions $l = 20$ cm. and $a = 1.1 - 1.2$ cm. Here, also

$$d = \frac{-a^2}{6l} + \frac{a}{3}; \qquad \begin{array}{l} d = 0.36 \text{ cm. for } a = 1.1 \text{ cm.} \\ d = 0.39 \text{ cm. for } a = 1.2 \text{ cm.} \end{array}$$

The cell used by Spence and Wild (72) was of annular construction, with 1 cm. between walls. This cell is analyzed according to *Case 2*, with $a \rightleftarrows l$. The cell was 10 cm. long ($a = 10, l = 1$).

$$d = \frac{l^3 - 8l^2 a + 24a^2 l}{96a^2} = 0.25$$

The short cell used by Howland and Noyes (39) also is analyzed according to *Case 2*, with $l = 1$ cm., $a = 1.1$ cm.

$$d = 0.18 \text{ cm.}$$

found by Spence and Wild (71, 72) and by Noyes *et al.* (36, 39) may come from an examination of the light sources used. For banded absorption experiments Noyes *et al.* used approximately monochromatic light, the main impurity being due to the pressure broadening of the 3130 A. group in the medium-pressure quartz-jacketed mercury arcs. Spence and Wild (71) used one of three filters to remove continuum radiation: Pyrex, potassium biphthalate, potassium chromate. Only the latter filter removes 3347 and 3660 A. radiation to a large extent, and only when this filter was used was a value of C_2H_6/CO greater than unity obtained. Biacetyl might decompose either by absorption of wave lengths 3347 and 3660 A. or by collision of the second kind with acetone (see above). The ratio C_2H_6/CO from acetone would be much more nearly unity than either with small percentage decomposition or in a flow system where the biacetyl is removed. This effect would be much more pronounced when the continuum light is removed, since then the 3660 A. radiation constitutes a much larger per cent of the total incident intensity.

The effect of increase of pressure on quantum yield and C_2H_6/CO (condition *a*) could not be explained on the basis of an excited acetone molecule of mean life 10^{-3} sec., a figure determined by Spence and Wild (72) from the above data. If the effect were due to deactivation of an excited molecule by collisions, such as we have in unimolecular reactions, a life-period of about 10^{-8} sec. would be expected. Spence and Wild consider that the widely different requirements for excited acetone molecules are met by the two excited states. The reactions of the excited state of life-period 10^{-3} sec. are given by reaction 2a (low pressure) and by reaction 13 (high pressure). The reactions of the long-life excited molecules are given in equations 14–17.

On the basis of a purely free-radical mechanism, the effect of increase of pressure on quantum yield and C_2H_6/CO is readily explained by the homogeneous reactions 6 or 7 (increased dissociation of acetyl leads to a smaller value of C_2H_6/CO); the recombination reaction 8; the decreasing importance of 11 with increase of pressure and the smaller increase of the importance of reactions 11 and 12.

The Spence and Wild theory (72) offers an explanation of the decrease of the C_2H_6/CO ratio with decreasing wave length that differs from the free-radical explanation. The latter theory considers the increased energy retained by the acetyl radicals after dissociation. As this energy increases (decreasing wave length) the spontaneous decomposition must increase, leading to smaller values of C_2H_6/CO. The Spence and Wild view is that the excess vibrational energy possessed by the acetyl radicals would be rapidly dissipated by collision at high pressures, but with decrease of pressure a growing proportion would decompose spontaneously. In consequence, the concentration of acetyl radicals, and therefore the value of C_2H_6/CO, would be less in the "continuum" region than in the "banded" region at low pressures.

As regards the effect of mixed light (condition *d*), the Spence and Wild view is that excited acetone molecules of long life produced in the banded region bring about the decomposition of acetyl radicals produced in the continuum (reaction 17).

On the basis of any theory the explanation of the changes of experimental quantities (ethane, carbon monoxide, etc.) with increasing polychromatic character of the light source must be very difficult at best. Certainly the products formed by irradiation with any two different mercury arcs, unfiltered, would be different if the spectral distributions were different. Because of this, and because of chromatic aberration effects, it is nearly impossible to account quantitatively for the effect of mixed light.

The results of the Spence and Wild mechanism, which includes the reactions 1, 7, 8, 9′, 10, 11′ as radical steps plus the excited molecule steps 2, 2a, 13–17 is given graphically by these authors, the relative importance of the two mechanisms being plotted as functions of acetone pressure.

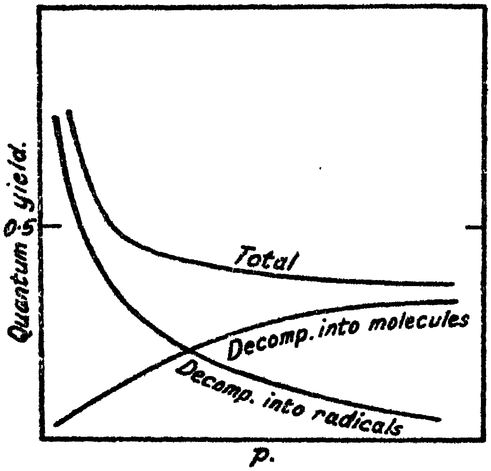


FIG. 7. Variation of quantum yield of photodecomposition of acetone with pressure in the banded region (qualitative). (Spence and Wild (72).)

The experimental data on the photolysis of acetone at 3130 A., the "banded" region, contradict one of the theories (free radical or excited acetone) to some extent, if the light used by Spence and Wild (71, 72) can be considered to consist of radiation between the limit 3050 A. and 3300 A. This point, as noted above, seems doubtful. As a result it seems that the excited-molecule theory requires more conclusive proof (the diameter effect) before it can be accepted. The data of Noyes *et al.* (36, 39) cannot be explained by the Spence and Wild theory.

One possible method of differentiation between the two theories, besides further studies on the diameter effect using nearly monochromatic 3130 A. radiation, would be to photolyze CH_3COCH_3 in the presence of CD_3COCD_3. Such studies would shed considerable information on the reactions by which ethane is formed. If this is formed by a non-radical mechanism under any conditions, C_2H_6 and C_2D_6 would be formed to a greater extent than would be expected upon the basis of a free-radical mechanism, which would predict, approximately:

$$C_2H_6:CH_3CD_3:C_2D_6 = 1:2:1$$

At pressures in the order of 150 mm. of acetone, or higher, the Spence and Wild mechanism would predict this ratio to be of the order 1:1:1, quite different from the predictions of the purely free-radical mechanism. Further, the acetone remaining after illumination should be a mixture of CH_3COCH_3, CH_3COCD_3, and CD_3COCD_3 if the free-radical mechanism is obeyed; i.e., if the low quantum yield of acetone decomposition is due to radical recombination.

Before discussing the effect of temperature upon the products of the photolysis of acetone, it might be well to consider some recent work of Feldman, Burton, Ricci, and Davis (24). The excited-molecule mechanism was presented by Spence and Wild to account for the results of the photolysis of acetone only in the "banded" region of absorption above about 3050 A. Below this wave length, that is, in the region of continuous absorption, these authors, in agreement with others, had considered that only the free-radical mechanism was necessary to explain the experimental data. As a result of their experiments on the flow photolysis of acetone using 2537 A. radiation, Feldman, Burton, Ricci, and Davis now consider that the Spence and Wild mechanism must be active in this region also. This conclusion was reached as a result of the formation of considerable quantities of ethane under conditions such that nearly all methyl radicals were expected to be removed by lead mirrors before these radicals could react to form ethane. If a quantum yield for carbon monoxide formation of about 0.2 is assumed for low pressures of acetone (2 mm.) at 2537 A., the data of Feldman, Burton, Ricci, and Davis indicate an absorbed intensity of about 2×10^{15} quanta/sec., that is, intensities about twenty times greater than those used by Herr and Noyes (36), and of the same magnitude as used by Benson and Forbes (10). The work of Benson and Forbes (10), the flow photolysis in the absence or presence of iodine, shows no ethane formed at temperatures up to 140°C. when iodine is present. This is opposed to the data of Feldman, Burton, Ricci, and Davis, who found large amounts of ethane formed in the presence of another radical remover, the lead mirror. To reconcile these different results would be rather difficult and will not be attempted, but a few obvious conclusions can be drawn. Feldman, Burton, Ricci, and Davis have considered that, following reaction 2, reactions 2a and 13 occur. To explain their results, reaction 13 would have to predominate at low pressures, where, according to the Spence and Wild mechanism, reaction 2a predominates, at least at 3130 A. Further, Feldman, Burton, Ricci, and Davis believe that in the presence of iodine, reaction 13 is followed by

$$CH_3COCH_3^{**} + I_2 \rightarrow CH_3I + CH_3COI \qquad (18)$$

To explain the complete absence of ethane in the work of Benson and Forbes, reaction 18 would have to have an efficiency of about 100 per cent; i.e., to have occurred so rapidly that reactions 14 and 16, the ethane-forming reactions, are relatively unimportant.

Although Feldman, Burton, Ricci, and Davis reported methyl radicals, carbon monoxide, and ethane to be their main products, other authors, Benson and Forbes (10), Herr and Noyes (36), and Spence and Wild (71), have all found biacetyl formed at 2537 A. Benson and Forbes (10) reported a $(CH_3CO)_2/CO$ ratio of 0.1–0.4 at 60°C., depending upon conditions, but apparently increasing with linear flow velocity. Thus, although Feldman, Burton, Ricci, and Davis reported no biacetyl, this must certainly have been formed in their experiments, especially at 0°C. If the analyses of Feldman, Burton, Ricci, and Davis for methyl, carbon monoxide, and ethane are assumed correct, the amount of biacetyl calculated from the material balance

$$\begin{aligned} N_{\text{acetone}} &= 2N_{C_2H_6} - N_{CO} \\ N_{\text{biacetyl}} &= N_{C_2H_6} - N_{CO} \\ \tfrac{1}{2}N_{CH_3} + n_{C_2H_6} &= N_{C_2H_6} \end{aligned} \tag{19}$$

$n_{C_2H_6}$ = cm.^{3} or mm.^{3} of ethane analyzed as such.

is always found to be a negative quantity of considerable magnitude at 0°C. and a positive quantity at other temperatures up to 100°C., with a few exceptions at 100°C. Further, although the average deviation from a correct material balance for all runs is only about 25 per cent, individual runs show errors of up to 85 per cent. If the ethane and carbon monoxide analyses are assumed to be correct, using equation 19, the error in the methyl determination appears to be enormous in the work of Feldman, Burton, Ricci, and Davis, the value of CH_3(missing)/CH_3(found) varying from 3.5 to − 8.0.

Aside from these considerations of the work of Feldman, Burton, Ricci, and Davis, the Spence and Wild mechanism extended to continuum absorption (2200–3000 A.) contradicts the work of Taylor and Rosenblum (74). Table 12 shows that in the presence of hydrogen the photolysis of acetone yields a value of $CH_4/CO \simeq 1$ at about 250°C. Since methane can presumably be formed only from the intermediate methyl radical, the above ratio requires a minimum of 50 per cent radical decomposition of acetone, allowing no formation of ethane from methyl. Further, at 300°C., $CH_4/CO \simeq 1.35$, here requiring a minimum of 68 per cent radical decomposition of acetone.

Considering the many questions that can be raised concerning the validity of the data of Feldman, Burton, Ricci, and Davis, it appears that this work cannot be considered of sufficient accuracy to warrant the discarding of the free-radical mechanism for the photolysis of acetone at 2537 A. in favor of an extended Spence and Wild mechanism.

As in the case of photolysis in the "banded" region, it is quite probable that if the Spence and Wild mechanism is correct for "continuum" photolysis, this could be determined by a thorough study of the photodecomposition of ethyl methyl ketone. Here again, the Spence and Wild mechanism would require that $C_2H_6:C_3H_8:C_4H_{10}$ be considerably different from 1:1:1 if the hydrocarbon can be formed in one step. Also, the photolysis of mixed normal and deutero acetones would yield data that would allow differentiation between the free-radical mechanism and the Spence and Wild mechanism, as will further determinations on the extent of methane and C_3- and C_4-hydrocarbon formation.

The effect of temperature upon the photolysis of acetone has been studied to a considerable extent. As previously mentioned, the formation of biacetyl becomes small at temperatures above 60–75°C. Gorin (30) has shown that the absence of biacetyl up to 100°C. may be due to the reaction

TABLE 6

Experiments in unfiltered light on the stability of the acetyl radical and the formation of acetyl iodide

(Gorin (30))

PRESSURE OF ACETONE	INITIAL PRESSURE OF IODINE	CH_3I	CH_3COI	CO	TEMPERATURE
mm.	*mm.*	*moles* × 10^6	*moles* × 10^6	*moles* × 10^6	°*C.*
165	4	21.7	9.3	1.62	100
487	2	43.0	4.0	2.67	100
430	4	29.4	8.8	5.28	115
156	2	36.8	4.6	5.82	115
517	4	14.6	7.0	4.96	128
169	2	41.6	3.4	15.02	130

TABLE 7

Photolysis of acetone, acetone–iodine mixtures at λ 2537 A.

$P_{acetone}$ = 12.0 ± 0.2 mm. (0°C.); $P_{acetone}/P_{iodine}$ = 54 ± 2

(Benson and Forbes (10))

RUN	TEMPERATURE	FLOW	CC. AT N.T.P.				IODIDES	$\frac{CH_4}{CO}$	$\frac{C_2H_6}{CO}$	$\frac{(CH_3CO)_2}{CO}$
			CO	CH_4	C_2H_6	$(CH_3CO)_2$				
	°C.	*cm./sec.*								
4	60.1	1.05	2.34	0.18	2.81	0.90		0.078	1.20	0.38
5*	60.0	?	2.11	0.061			?	0.029		
6	59.9	0.79	2.48	0.21	3.06	0.73		0.083	1.23	0.29
7*	59.9	0.53	2.09	0.036			5.42	0.017		
8	100.3	0.97	3.45	0.23	3.49	0.10		0.067	1.01	0.029
9*	100.5	0.88	1.38 (?)	0.021			3.60	0.015		
10	142	0.62	3.68	0.37	3.35	0.00		0.100	0.911	0.00
11*	142	0.67	2.03	0.029			4.26	0.014		
12	60.2	0.54	2.00	0.14	2.20	0.22		0.068	1.10	0.11
13*	60.2	0.65	1.68	0.022			4.46	0.013		
14*	102.0	0.63	1.84	0.038			4.16	0.020		
15	60.5	0.96	2.02	0.13	2.37	0.44		0.066	1.17	0.22

* Runs made with iodine present.
Time of runs = 90 min.

$$CH_3 + CH_3CO \rightarrow C_2H_6 + CO \qquad (10)$$

rather than to the increased importance of

$$CH_3CO \rightarrow CH_3 + CO \qquad (6)$$

although, as mentioned previously, reaction 6 is probably more important at higher temperatures than Gorin indicates. Some of the experimental details are missing from the paper of Gorin (30), so the sufficiency of the iodine present to react with all the methyl and acetyl formed might be questioned. However, in an earlier paper Gorin (28) stated that 0.2 mm. iodine was sufficient to react with all radicals formed. (For experimental details of Gorin's work see reference 29.) Using this, we would expect the 0.5 mm. iodine, the minimum amount reported used (30), to be sufficient. If reaction 6 became very important at 100°C., considerable amounts of carbon monoxide would be expected at this temperature, even in the presence of iodine. Gorin (30) and Benson and Forbes (10) found less carbon monoxide formed in the presence of iodine than in the absence, even up to temperatures of 140°C. (see tables 6 and 7). This is not opposed to the increasing importance of the thermal reaction (6).

The stability of the acetyl radical has also been demonstrated by Feldman, Ricci, and Burton (25) by the use of the mirror technique. Acetone was irradiated at 45°C. with 2537 A. light, and after removal of methyl radicals by a first mirror, heating the flow tube gently caused the disappearance of a second mirror, probably by methyl radicals formed from the thermal decomposition of CH_3CO. The reduction in the relative amount of biacetyl above 60–75°C. has been attributed (30) to the reaction

$$CH_3 + CH_3CO \rightarrow C_2H_6 + CO \tag{10}$$

rather than to the instability of the acetyl radical. However, as explained above, biacetyl has been found (6) up to 110°C., so that pressure and intensity at this temperature, plus other reactions, determine whether or not biacetyl is formed.

Taylor and Jungers (73) suggested another reaction to explain the absence of biacetyl above about 75°C. and also partly to account for the low over-all yield of acetone decomposed at lower temperatures.

$$CH_3 + (CH_3CO)_2 \rightarrow CH_3COCH_3 + CH_3CO \tag{20}$$

Herr, Matheson, and Walters (35) photolyzed acetone in the presence of $(CD_3CO)_2$ and found no deuterated ethane, such as CH_3CD_3. Since this latter product would be expected if reaction 20 occurred to an appreciable extent, its absence is strong evidence that reaction 20 does not occur, at least at the temperatures (below 60–75°C.) where biacetyl formation is favored.

Anderson and Rollefson (6) have actually reported the formation of biacetyl at temperatures as high as 110°C. (see table 8).

That the photochemical decomposition of acetone does not follow a chain process has been demonstrated by many workers. Leermakers (42) showed that the quantum yield for carbon monoxide formation does not exceed 1.3 even up to 400°C. Winkler (76) obtained essentially the same results as Leermakers, and also reported the formation of compounds not found when the photolysis occurs at room temperature. Thus, for 50 per cent decomposition of acetone at 60°C., Winkler found approximately 9 per cent methane, while at 400°C. 32 per cent methane was found, 5 per cent hydrogen, and some unsaturated compounds.

In the presence of nitric oxide, Mitchell and Hinshelwood (47) confirmed the above, finding $\Phi_{\text{acetone}} = 1.1$ as the average of eight runs at 300°C. The only difference caused by the presence of nitric oxide was the formation of carbon dioxide.

At 75°C. Herr and Noyes found the quantum yield for carbon monoxide formation to rise from about 0.8 to 0.9 as the pressure of acetone was increased from about 5 to 150 mm., both at 3130 A. and 2537 A. at absorbed intensities of 5×10^{13} quanta per second (figures 2 and 3). However, the quantity was found to decrease with increasing intensity at 50 mm. pressure (36).

TABLE 8

Photolysis of acetone at 2537 A.

Reaction cell 5.0 cm. in diameter, 5.5 cm. long; volume 113.5 cc. with 3.3 per cent dead space

(Anderson and Rollefson (61))

TEMPERATURE	INITIAL $(CH_3)_2CO$	PRODUCTS IN MOLES $\times 10^5$			
		CO	CH_4	C_2H_6	$(CH_3CO)_2$
°C.	*moles* $\times 10^5$				
3	24.39	0.80	0.09	2.85	2.20
3.5	24.75	1.10	0.17	2.93	2.18
22	25.82	0.74	0.08	2.69	2.01
22.5	25.33	0.87	0.12	2.58	1.76
26.5	25.23	0.72	0.09	2.70	2.04
64.5	24.95	1.68	0.10	3.16	1.56
65.5	24.26	1.49	0.13	2.74	1.40
82	24.18	2.99	0.17	4.09	1.21
88	24.55	3.66	0.19	4.57	1.19
98	24.71	4.51	0.15	5.30	0.86
102	24.63	4.34	0.16	4.70	0.59
107.5	24.47	2.53	0.23	2.67	0.26
109.5	25.32	3.72	0.22	3.91	0.24
110	24.63	4.46	0.15	4.62	0.27
114.5	25.33	4.46	0.18	4.48	0.11
120.5	24.59	5.80	0.20	5.70	0
133	24.02	3.63	0.21	3.57	0
140	25.20	5.45	0.40	5.29	0

To get concentrations in moles per liter, multiply moles by 8.8.

That the quantum yields of carbon monoxide formation or acetone decomposition appear to approach unity at about 125°C. is shown in the work of Herr and Noyes (36).

The way in which the quantum yields of acetone decomposed or carbon monoxide formed decrease with intensity has not been studied very carefully. Thus, although these quantities approach unity at about 125°C., Akeroyd and Norrish (1) found values considerably less than unity when acetone is photolyzed using the full intensity of a mercury arc (see table 9).

Akeroyd and Norrish also reported the distribution of products shown in table 10 for the photolysis at 350°C.

At high intensities at 860 mm. acetone pressure Spence and Wild (71) found considerable amounts of methane even at 60°C. (see table 11).

Of course, runs at 860 mm., as mentioned by Spence and Wild, cannot be compared with runs say at 200 mm. or lower acetone·pressure. Hence, these data are of use only in indicating that even at 60°C. methane may be formed at the lower pressures.

TABLE 9

Photolysis of acetone with full intensity of a mercury arc

P_{acetone} = 70 mm. when reduced to 100°C.

(Akeroyd and Norrish (1))

Temperature, °C	63	100	200	300	396
Mean Φ_{acetone}	0.3	0.47	0.5	0.65	0.76

TABLE 10

Photolysis of acetone at 350°C.

Time of illumination = 284 min.

(Akeroyd and Norrish (1))

Product	CO	CH_4	C_2H_6	Unsaturated	H_2
cc. at N.T.P.	1.0	0.8	0.4	0.15	0.1

TABLE 11

Photolysis of acetone at high intensities

(Spence and Wild (71))

NO.	RADIATION	TIME	SCREEN TRANSMISSION	GASEOUS PRODUCTS			
				Volume at N.T.P.	CO	C_2H_6	CH_4
		hr.	*per cent*	*cc.*	*per cent*	*per cent*	*per cent*
24	Full arc	9.5	100	13.52	45.0	41.4	13.6
26	Continuum	24	100	4.15	48.3	33.3	18.4
27	Continuum	20	100	3.27	47.6	32.6	19.8
30	Continuum	24	29	1.54	52.6	19.6	27.8
28	Continuum	15	22	0.73	55.0	18.9	26.0
31	None	24		0.00			

Taylor and Rosenblum (74) have studied the effect of temperature upon the photolysis of acetone in the presence and absence of hydrogen (table 12). They found no increase in Φ_{CO} from 160° to 300°C. in the absence or presence of hydrogen, but the presence of hydrogen caused an increase in the methane yield and a decrease in the ethane yield.

The effect of hydrogen upon the ratios CH_4/CO and C_2H_6/CO is shown in table 13.

The values of C_2H_6/CO being greater than unity below 160°C. would suggest that below this temperature biacetyl is formed. From the lack of carbon balance, determined by $(C_2H_6 + \frac{1}{2}CH_4)/CO$, it appears, as stated by Taylor and Rosenblum, that above 70°C. the hydrocarbons start to polymerize, this factor becoming quite important at 200°C.

TABLE 12

The photolysis of acetone in the presence and absence of hydrogen

Reaction volume, 131.4 cc.; full mercury arc; light below 2000 A. removed

(Taylor and Rosenblum (74))

TEMPERATURE	MOLES OF REACTANT $\times 10^4$		MOLES OF PRODUCTS $\times 10^4$					
	$(CH_3)_2CO$	H_2	$(CH_3)_2CO$ By difference	$(CH_3)_2CO$ Measured	CH_4	CO	H_2	C—C
°C.								
71	2.52		0.81	0.72	0.109	1.305		1.524
71	2.57	12.85	0.87	0.82	0.101	1.383	12.79	1.512
160	2.52		1.07	1.01	0.079	1.377		1.352
160	2.55	12.74	0.90	0.86	0.423	1.582	12.61	1.353
200	2.49	12.44	0.88	0.83	0.933	1.592	12.11	0.949
250	2.52		0.76	0.71	0.463	1.753		1.213
250	2.50	12.49	0.95	0.89	1.57	1.51	11.96	0.482
300	2.49		0.86	0.71	0.802	1.513		0.863*
300	2.54	12.69	0.77	0.72	2.327	1.728	11.69	0.378

* This value may be high by 8 per cent, which would yield for the final acetone values approximately 0.92 by difference and approximately 0.77 measured.

TABLE 13

Effect of hydrogen upon the CH_4/CO *and* C_2H_6/CO *ratios*

(Taylor and Rosenblum (74))

TEMPERATURE	C_2H_6/CO		CH_4/CO		$\frac{2(-\Delta H_2)}{CO}$	$\frac{C_2H_6 + 1/2\ CH_4}{CO}$	
	H_2 absent	H_2 present	H_2 absent	H_2 present		H_2 absent	H_2 present
°C.							
71	1.17	1.09	0.083	0.073	0.087	1.21	1.13
160	0.982	0.855	0.057	0.268	0.164	1.01	0.99
200	(0.86)*	0.595	(0.11)*	0.585	0.414	(0.91)*	0.89
250	0.692	0.319	0.264	1.04	0.702	0.82	0.84
300	0.570	0.219	0.530	1.35	1.16	0.83	0.89

* Interpolated values.

In the presence of hydrogen the formation of methane can easily be explained by reaction 21:

$$CH_3 + H_2 \rightarrow CH_4 + H \qquad (21)$$

Taylor and Rosenblum determined an energy of activation of 11 ± 2 kg.-cal. for this reaction.

The method of formation of methane in the absence of hydrogen has generally been taken to be that listed by Rice and Herzfeld (65):

$$CH_3 + CH_3COCH_3 \rightarrow CH_4 + CH_3COCH_2 \qquad (22)$$

The energy of activation of this reaction has been given as about 15 kg.-cal. (14, 65).

A reaction to be expected as a consequence of reaction 22 would lead to the formation of ethyl methyl ketone as well as ketene.

$$CH_3COCH_2 + CH_3 \rightarrow CH_3COCH_2CH_3 \qquad (23)$$

$$CH_3COCH_2 \rightarrow CH_2CO + CH_3 \qquad (24)$$

Reaction 23 has been assigned an energy of activation of about 8 kg.-cal. (14), while the 48 kg.-cal. energy of activation of reaction 24 (14) suggests that the formation of ketene will be unimportant except at temperatures above 400°C.

Saunders and Taylor (69) have carried out a very extensive study on the effects of mercury and dimethylmercury upon the photolysis of acetone at temperatures from 100° to 275°C. The data given by these authors are so extensive that they are best represented graphically. In general, below about 100°C. it appears that mercury and even dimethylmercury have very little effect upon the products. However, above about 150°C. the presence of these substances is seen from figure 8 to cause a decrease in the amounts of carbon monoxide and ethane formed and an increase in the amount of methane formed.

Saunders and Taylor also found the carbon deficiency, or the amount of hydrocarbons with more than three carbon atoms, to be a function of mercury and dimethylmercury as well as of temperature (see figure 9). Thus, these substances have an effect upon the polymerization of the lower hydrocarbons.

Most of the reactions considered by Saunders and Taylor for the photolysis of acetone in the absence of mercury and dimethylmercury are those already considered, namely:

	E_{act}	
$CH_3COCH_3 + h\nu \rightarrow CH_3 + CH_3CO$		(1)
$CH_3 + CH_3CO \rightarrow CH_3COCH_3$	0–8	(8)
$CH_3 + CH_3CO \rightarrow C_2H_6 + CO$	6	(10)
$CH_3CO + (CH_3COCH_3) \rightarrow CH_3 + CO + (CH_3COCH_3)$	18	(7)
$CH_3 + CH_3COCH_3 \rightarrow CH_4 + CH_2COCH_3$	16	(22)
$2CH_2COCH_3 \rightarrow (CH_3COCH_2)_2$		(25)
$CH_3CO + CH_2COCH_3 \rightarrow CH_3COCH_2COCH_3$		(26)
$CH_3 + CH_2COCH_3 \rightarrow C_2H_5COCH_3$	8	(23)
$CH_3 + C_2H_5COCH_3 \rightarrow CH_4 + C_2H_4COCH_3$	<16	(27)
$C_2H_5COCH_3 + h\nu \rightarrow (CO + C_2H_6 + C_3H_8 + C_4H_{10})$		(28)

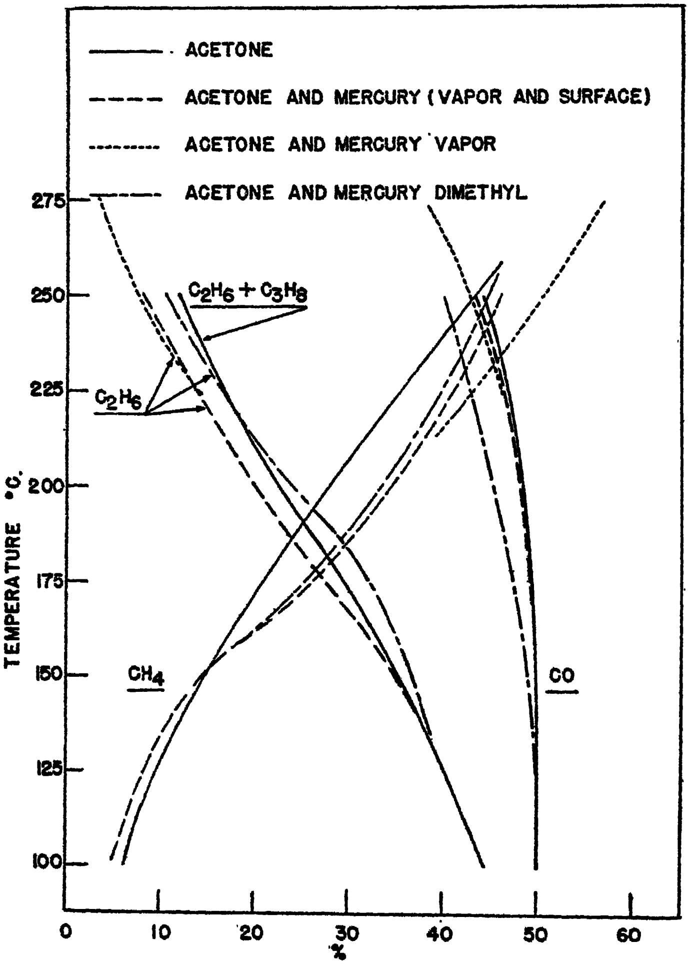


FIG. 8. Effect of temperature on the products of the decomposition of acetone. (Saunders and Taylor (69).)

In the presence of mercury, Saunders and Taylor have presented the following additional equations:

	E_{act}	
$CH_3 + Hg \rightarrow HgCH_3$		(29)
$CH_3 + HgCH_3 \rightarrow Hg(CH_3)_2$		(30)
$CH_3 + HgCH_3 \rightarrow CH_4 + HgCH_2$	13	(31)
$CH_3 + HgCH_3 \rightarrow C_2H_6 + Hg$		(32)

Additional reactions suggested for dimethylmercury are:

$$CH_3 + Hg(CH_3)_2 \rightarrow CH_4 + Hg(CH_3)CH_2 \qquad (33)$$

$$CH_3 + Hg(CH_3)_2 \rightarrow C_2H_6 + HgCH_3 \qquad (34)$$

Saunders and Taylor have used equation 10 to explain the presence of ethane rather than equation 6 or 7. This use may be justified, since the pressures used were about 25 to 200 mm. at relatively high intensities, but the thermal reaction (reaction 6) certainly is very important above 150°C. At the lower pressures equation 6 or 7 also should be considered.

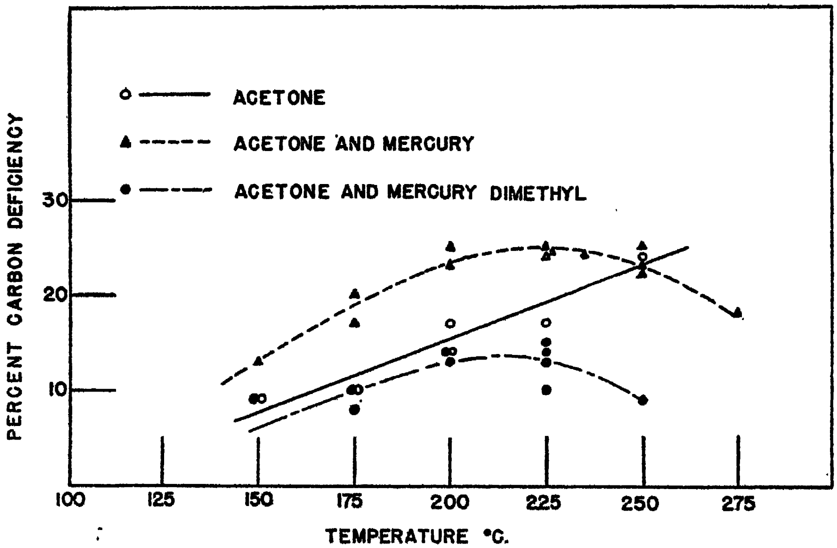


FIG. 9. Per cent carbon deficiency as a function of temperature. (Saunders and Taylor (69).)

Two conclusions drawn by Saunders and Taylor (69) are of interest, namely:

$$\frac{CH_4}{CO} \propto (P_{\text{acetone}})^{2/3} \tag{35}$$

and

$$\frac{C_2H_6}{CO} \propto \left(\frac{1}{P_{\text{acetone}}}\right)^{1/3} \tag{36}$$

A few last points on the photolysis of acetone at higher pressures need mentioning. Roof and Blacet (67) have pointed out that at temperatures not much above 100°C. the acetyl radical decomposes nearly to the extent of 100 per cent according to reaction 6 or 7. This is not in disagreement with previous statements that at 100°C. 60–90 per cent of the acetyl radicals decompose. Thus, the formation of ethane can result nearly solely from the combination of two methyl radicals. At high pressures, either of acetone or of foreign gases, the diffusion of methyl radicals to the walls would be very slow, and, hence, if wall formation of ethane were the only way in which this compound could be formed, under these conditions we would expect homogeneous reactions such as 22 to predominate, the rate of ethane formation being low. However, Grahame and

Rollefson (32) have obtained evidence that ethane is formed by the rather accurately bimolecular reaction:

$$CH_3 + CH_3 \rightarrow C_2H_6 \qquad (9')$$

At high temperatures and pressures ethane possibly results solely according to equation 9′.

Studies by Danby and Hinshelwood (17) on the polymerization of olefins induced by free radicals at 300°C. have led these authors to conclude that acetone

TABLE 14

Polymerization of olefins by free radicals at 300°C.

(Danby and Hinshelwood (17))

	I_a (ARBITRARY UNITS)
50 mm. CH_3CHO	9.9
25 mm. CH_3COCH_3	5.1
50 mm. CH_3COCH_3	8.4
25 mm. $C_2H_5COC_2H_5$	5.8

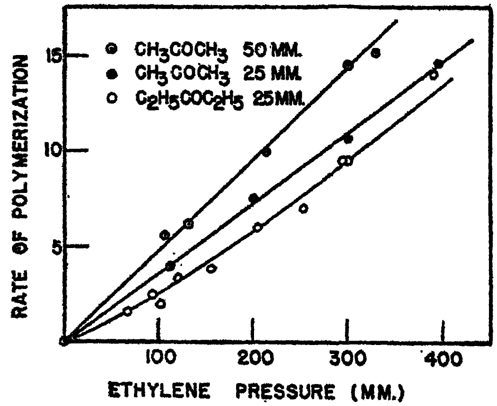


FIG. 10. Polymerization of ethylene induced by photolysis of dimethyl and diethy ketones. (Danby and Hinshelwood (17).)

is only 69 per cent as effective in producing free radicals at this temperature as is acetaldehyde (table 14; figure 10). In this work the full radiation of the mercury arc was used. The rate of the induced polymerization, σ, was calculated from the equation

$$\sigma = nk_{38}[C_2H_4]\sqrt{\frac{k_1 I_a}{k_{9'}}} \qquad (37)$$

where

$$CH_3COCH_3 + h\nu \rightarrow CH_3 + CH_3CO \qquad k_1 \qquad (1)$$

$$CH_3 + CH_3 \rightarrow C_2H_6 \qquad k_{9'} \qquad (9')$$

$$CH_3 + C_2H_4 \rightarrow R_1 \qquad k_{38} \qquad (38)$$

Since polychromatic light was used in these experiments, the ratio of the rates of polymerization induced by acetone and acetaldehyde is necessarily some sort of average over wave lengths from about 1800 A. (the lower transmission wave length of quartz) to about 3300 A.

Grahame and Rollefson (32) have made similar comparisons on the relative rates of radical production by acetone and acetaldehyde during studies on the high-temperature chain decomposition of acetaldehyde (table 15).

The value of 0.70 obtained for acetone at 3130 A. by Grahame and Rollefson is in good absolute agreement with that obtained by Gorin (30) for the photolysis of acetone at 90°C. in the presence of iodine (see table 16).

TABLE 15

Relative quantum efficiencies for the primary photochemical processes in the dissociation of acetone and acetaldehyde into free radicals at elevated temperatures

(Grahame and Rollefson (32))

λ	ACETALDEHYDE	ACETONE
A.		
3130	1.00	0.70
2652	0.55	0.92

TABLE 16

Photolysis of acetone, acetone–iodine mixtures

Experiments in filtered light; λ = 3130 A.

(Gorin (30))

EXPT.	PRESSURE OF ACETONE	PRESSURE OF IODINE	TEMPERATURE	CH_3I	CO	TIME	CH_3I
	mm.	*mm.*	°C.	*moles* × 10^6	*moles* × 10^6	*min.*	
1	852	1	90	5.67		120	0.85
2	852	0	80		1.13	120	
3	1265	0.5	90	8.76		120	0.78
4	1265	0	80		1.90	120	

That hydrocarbons at least as high as butane can be formed in the high-temperature photolysis of acetone has been demonstrated by Allen (2). Experiments were run using an unfiltered, high-intensity mercury arc, with about 50 per cent of the acetone decomposed. Besides methane, Podbelniak distillation analyses showed the presence of propylene, propane, and butane (normal and iso).

An accurate determination of the mechanism of the acetone photolysis above about 300°C. is rather difficult, owing to the apparent catalytic effects of the walls. Thus, Allen (2) considers that methyl radicals and acetone can react at the wall to give methane and acetonyl. The latter does not seem to dimerize on the wall, but only reacts with a methyl radical to form ethyl methyl ketone.

The higher hydrocarbons are considered to be formed by reactions such as the following:

$$CH_3 + C_2H_6 \rightarrow CH_4 + C_2H_5 \qquad (39)$$

$$C_2H_5 + CH_3 \rightarrow C_3H_8 \qquad (40)$$

$$CH_3 + C_3H_8 \rightarrow CH_3CHCH_3 \text{ or } CH_3CH_2CH_2 + CH_4 \qquad (41)$$

$$C_3H_7 + CH_3 \rightarrow C_4H_{10} \text{ (normal and iso)} \qquad (42)$$

Since the acetone photolysis can become so complicated above 300°C., studies at these temperatures are probably of little help in determining the nature of the over-all scheme of decomposition.

Nothing has been said so far concerning the photolysis of acetone at wave lengths below 2000 A. Manning (44) found that wave lengths of 1850–2000 A. were effective in causing the decomposition of acetone. At acetone pressures around 90 mm. the products were found to be carbon monoxide, hydrogen, methane, ethylene, and ethane. The percentage of these products was found to vary considerably with time, probably owing to secondary reactions. The experiments of Manning were performed in reaction vessels containing fluorite windows which transmit light below 1850 A. and, hence, were complicated by mercury resonance radiation (1849 A.) effects. No mechanism for the decomposition was proposed by Manning, but some qualitative results may be stated: (*1*) the C—H bonds are broken to a considerable extent, hydrogen forming about 63 per cent of the products for complete decomposition of acetone at pressures of about 0.025 mm.; (*2*) small amounts of carbon dioxide are formed at low pressures; (*3*) either the methyl radicals, if formed, react mainly to give methane, or if ethane or ethylene is formed, these products react further and polymerize; (*4*) at higher pressures (90 mm.), secondary reactions such as $H_2 + C_2H_4 \rightarrow C_2H_6$ occur.

Howe and Noyes (37) studied the photolysis of acetone at about 1900 A., using a quartz cell to remove light of wave length much below this value. The effect of mercury resonance radiation was also determined. At low pressures (2–8 mm. of acetone) carbon monoxide was found to constitute 50 per cent (within experimental error) of the reaction products. In agreement with Norrish, Crone, and Saltmarsh (53), Howe and Noyes (37) found the determination of ethane quite difficult, owing to solution in undecomposed acetone. Hence, the latter authors have based the quantum yield of acetone decomposed upon the quantum yield of carbon monoxide formed according to the over-all reaction:

$$(CH_3)_2CO \rightarrow C_2H_6 + CO \qquad (43)$$

At pressures of about 2 mm. the quantum yield was found to be about 0.6, decreasing slightly with pressure. An extrapolation to zero pressure would not necessarily appear to indicate an over-all quantum yield of unity. Hence, if the primary process is

$$CH_3COCH_3 + h\nu \rightarrow CH_3 + CH_3CO \qquad (1)$$

the recombination reaction

$$CH_3CO + CH_3 \rightarrow CH_3COCH_3 \tag{8}$$

would appear to be important.

As mentioned previously, the extent of the reaction

$$CH_3CO^* \rightarrow CH_3 + CO \tag{5}$$

is about 50 per cent at 1900 A. which, if the primary process 1 had an efficiency of about unity, would explain the value of $\Phi_{CO} = 0.5$ obtained by Howe and Noyes at about 50 mm. acetone pressure. At this pressure reactions 1, 8, and 5 appear sufficient to explain the data for the photolysis of acetone with light of about 1900 A. The presence or absence of biacetyl was not proved, since the absence of fluorescence may not indicate the absence of biacetyl at 1900 A. and since only Φ_{CO} was determined, except in a few runs. Assuming no complicating factors, reactions 8 and 5 seem to be equally probably at 50 mm. pressure, since no biacetyl was found. At lower pressures, the reactions

$$CH_3CO \rightarrow CH_3 + CO \tag{6}$$

or

$$CH_3CO + M \rightarrow CH_3 + CO + M \tag{7}$$

or

$$CH_3 + CH_3CO \rightarrow C_2H_6 + CO \tag{10}$$

as well as

$$CH_3 \xrightarrow{\text{walls}} \tfrac{1}{2}C_2H_6 \tag{9}$$

or

$$CH_3 + CH_3 \rightarrow C_2H_6 \tag{9$'$}$$

appear to enter into the scheme.

The addition of foreign gases (table 17) has the predicted effect of increasing the over-all yield for acetone decomposition. Specifically, reaction 6 or 7 becomes more important and 8 less important.

The presence of mercury effects an increase in the over-all quantum yield, probably owing to the formation of substances such as $HgCH_3$ by deactivation of the mercury (see table 18).

From the effect of added gases on the over-all quantum yield of acetone decomposed, it would appear that if the efficiency of the primary process (reaction 1) is not unity at 1900 A., it is certainly very close to it.

A. Summary of reactions

I. Free-radical reactions

$$CH_3COCH_3 + h\nu \rightarrow CH_3CO + CH_3 \tag{1}$$

Necessary to explain biacetyl formation.

$CH_3COCH_3 + h\nu \rightarrow CH_3COCH_3^*$ (2)

Necessary in the Spence and Wild mechanism but unnecessary in the free-radical mechanism.

TABLE 17

Effect of added gases on quantum yield of acetone decomposition

(Howe and Noyes (37))

ACETONE PRESSURE	FOREIGN GAS PRESSURE	QUANTUM YIELD	QUANTUM YIELD WITHOUT ADDED GAS	ADDED GAS
mm.	*mm.*			
1.82	9.8	0.81	0.61	CO_2
1.81	9.8	0.85	0.61	
1.72	18.4	0.88	0.62	
2.12	2.23	0.69	0.61	N_2
2.12	3.78	0.81	0.61	
2.34	7.38	0.74	0.61	C_2H_6
2.32	7.13	0.75	0.61	
2.30	12.30	0.80	0.61	

TABLE 18

Effect of Hg(6^1P_1) *on quantum yield for acetone decomposition*

(Howe and Noyes (37))

ACETONE PRESSURE	APPARENT QUANTUM YIELD	CALCULATED QUANTUM YIELD IN ABSENCE OF Hg RESONANCE RADIATION	REMARKS
mm.			
1.51	0.73	0.62	Resonance radiation present
0.115	0.72	0.65	
0.124	0.88	0.65	Trap in CO_2-$CHCl_3$
0.134	0.81	0.65	
20.4	0.53		Resonance radiation present
20.4	0.53		Resonance radiation removed
9.9	0.55	0.55	Resonance radiation present

$CH_3COCH_3 + h\nu \rightarrow C_2H_6 + CO$ (3)

No data to support this as being of more than minor importance.

$CH_3COCH_3 + h\nu \rightarrow 2CH_3 + CO$ (4)

Equivalent to reaction 1 followed by reaction 5.

$CH_3CO^* \rightarrow CH_3 + CO$ (5)

The use of this following reaction 1 makes it possible to use one primary process instead of the two of reaction 1 plus reaction 4. Here CH_3CO is considered to dissociate spontaneously.

$$CH_3CO \rightarrow CH_3 + CO \qquad (6)$$

The lack of pressure dependence of the thermal reaction above 50–100 mm. allows for the use of a unimolecular reaction.

$$CH_3CO + M \rightarrow CH_3 + CO + M \qquad (7)$$

Below 25 mm. pressure, the thermal decomposition of CH_3CO is pressure dependent.

$$CH_3CO + M \rightarrow CH_3CO^* + M \qquad (7a)$$

$$CH_3CO^* \rightarrow CH_3 + CO \qquad (7b)$$

$$CH_3CO^* + M \rightarrow CH_3CO + M \qquad (7c)$$

The three equations represent the reaction involved in the thermal decomposition of CH_3CO.

$$CH_3CO + CH_3 \rightarrow CH_3COCH_3 \qquad (8)$$

The low quantum yield of acetone decomposed after a primary decomposition process of nearly unit efficiency requires the reformation of acetone from its decomposition products.

$$2CH_3 \rightarrow C_2H_6 \qquad (9')$$

The formation of ethane has been found to be bimolecular at high pressures. Where specified, this equation represents the formation of ethane without consideration of the mode of formation (homogeneous or heterogeneous).

$$CH_3 \xrightarrow{\text{wall}} \tfrac{1}{2}C_2H_6 \qquad (9)$$

Necessary to explain the increase of $\Phi_{C_2H_6}$ with decreasing pressure at low pressures.

$$CH_3 + CH_3CO \rightarrow C_2H_6 + CO \qquad (10)$$

Arguments for and against this reaction.

$$2CH_3CO \rightarrow (CH_3CO)_2 \qquad (11')$$

The formation of $(CH_3CO)_2$ has been found to occur in a manner similar to the homogeneous formation of ethane (reaction 9′) at high pressures.

$$CH_3CO \xrightarrow{\text{wall}} \tfrac{1}{2}(CH_3CO)_2 \qquad (11)$$

Necessary to explain the increase of $\Phi_{(CH_3CO)_2}$ with decreasing pressure at low pressures.

$2CH_3CO + M \rightarrow (CH_3CO)_2 + M$ rate $= k_{12}(Ac)^2M^y$ (12)

With the exponent $y = 0.25$ the data of Spence and Wild agree with this equation. More thorough considerations might show this reaction to occur simultaneously with reaction 11′ at high pressures.

$CH_3 + (CH_3CO)_2 \rightarrow CH_3COCH_3 + CH_3CO$ (20)

This reaction has been shown not to occur.

$CH_3 + H_2 \rightarrow CH_4 + H$ (21)

Formation of large amounts of methane in the photolysis of acetone in the presence of hydrogen makes this reaction necessary.

$CH_3 + CH_3COCH_3 \rightarrow CH_4 + CH_3COCH_2$ (22)

One of the few ways in which methane can be formed from methyl radicals in the presence of acetone.

$CH_3 + CH_3COCH_2 \rightarrow CH_3COCH_2CH_3$ (23)

Following reaction 22, the presence of these two radicals indicates ethyl methyl ketone a logical product, although it has not been detected analytically owing to experimental difficulties.

$CH_3COCH_2 \rightarrow CH_2CO + CH_3$ (24)

This reaction is probably unimportant except at temperatures above 400°C.

$2CH_3COCH_2 \rightarrow (CH_3COCH_2)_2$ (25)

Logical following reaction 22, but biacetonyl has not been detected, owing to experimental difficulties.

$CH_3CO + CH_3COCH_2 \rightarrow CH_3COCH_2COCH_3$ (26)

This reaction is probably quite unimportant, since the acetyl radical is quite unstable above 125°C., a temperature at which acetonyl is not formed in large quantities except at high acetone pressures.

$CH_3 + CH_3COC_2H_5 \rightarrow CH_4 + CH_3COC_2H_4$ (27)

Probably important in the photolysis of ethyl methyl ketone, but unimportant in the photolysis of acetone, the concentration of ethyl methyl ketone being very low in this case except for large percentages of decomposition.

$$CH_3COC_2H_5 + h\nu \rightarrow (CO + C_2H_6 + C_3H_8 + C_4H_{10}) \tag{28}$$

The ethyl methyl ketone photolysis is unimportant in the acetone photolysis except when large percentages of acetone are decomposed at elevated temperatures.

$$CH_3 + Hg \rightarrow Hg(CH_3) \tag{29}$$

$$CH_3 + Hg(CH_3) \rightarrow Hg(CH_3)_2 \tag{30}$$

$$CH_3 + Hg(CH_3) \rightarrow CH_4 + Hg(CH_2) \tag{31}$$

$$CH_3 + Hg(CH_3) \rightarrow C_2H_6 + Hg \tag{32}$$

$$CH_3 + Hg(CH_3)_2 \rightarrow CH_4 + Hg(CH_3)(CH_2) \tag{33}$$

$$CH_3 + Hg(CH_3)_2 \rightarrow C_2H_6 + Hg(CH_3) \tag{34}$$

Reactions 29–34 can occur when acetone is photolyzed in the presence of mercury and dimethylmercury.

$$CH_3 + C_2H_4 \rightarrow R_1 \tag{38}$$

The polymerization of olefins is induced by the photolysis of acetone.

$$CH_3 + C_2H_6 \rightarrow CH_4 + C_2H_5 \tag{39}$$

$$C_2H_5 + CH_3 \rightarrow C_3H_8 \tag{40}$$

$$CH_3 + C_3H_8 \rightarrow CH_3CHCH_3 \text{ (or } CH_3CH_2CH_2) + CH_4 \tag{41}$$

$$C_3H_7 + CH_3 \rightarrow C_4H_{10} \text{ (normal and iso)} \tag{42}$$

Reactions 39–42 are used to explain the high-temperature products of the photolysis of acetone.

II. Additional reactions of the Spence and Wild mechanism for the "banded" region near 3100 A.

$$CH_3COCH_3 \xrightarrow[\substack{\text{at low}\\ \text{pressures}}]{10^{-8}\text{ sec.}} CH_3 + CH_3CO \qquad \text{(2a}$$

At low pressure the Spence and Wild mechanism predicts mainly radical formation.

$$CH_3COCH_3^* + M \rightarrow CH_3COCH_3^{**} + M \tag{13}$$

At higher pressures, a second excited state is formed, this being used to explain the fluorescence of acetone and, with reactions 14, 15, and 16, to explain the decrease of C_2H_6/CO as the cell size decreases.

$$CH_3COCH_3^{**} \xrightarrow{10^{-2}\text{ sec.}} C_2H_6 + CO \tag{14}$$

The time factor was determined partly from the data on the decrease of C_2H_6/CO with cell size and from fluorescence data of biacetyl. A moderate and high pressure reaction.

$$CH_3COCH_3^{**} \rightarrow CH_3COCH_3 \qquad (15)$$

Deactivation by fluorescence or collision.

$$CH_3COCH_3^{**} \xrightarrow{\text{wall}} C_2H_6 + CO \qquad (16)$$

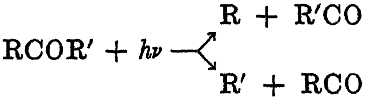
Since wall effects are quite small at 100 mm. pressure, this reaction should not be very important in this pressure range.

$$CH_3COCH_3^{**} + CH_3CO \rightarrow CH_3COCH_3 + CH_3 + CO \qquad (17)$$

This reaction is used to explain the decrease in C_2H_6/CO with increasing intensity at high intensities and at moderate to high pressures of acetone ($\geqslant$100 mm.)

III. ETHYL METHYL KETONE

The photolysis of the simple aliphatic ketones other than acetone has not been investigated to any great extent. Ethyl methyl ketone offers very interesting possibilities with regard to elucidating the mechanism of the photodecomposition of those ketones which undergo the Norrish Type I decomposition, namely



$$RCOR' + h\nu \rightarrow \begin{cases} R + R'CO \\ R' + RCO \end{cases} \qquad \text{Type I} \quad (1)$$

Acetone, of course, appears to decompose to a large extent, if not wholly, according to this Type I mechanism. If acetone decomposes to ethane and carbon monoxide without the intermediate formation of radicals, by analogy we would expect ethyl methyl ketone to do the same, at least after the absorption of light near 3100 A. Since the absorption spectrum of ethyl methyl ketone shows no structure around 3100 A. (22), it might not be proper to compare the results of the photolysis of acetone and ethyl methyl ketone in this wave-length region. Such a comparison would probably be proper if indisputable evidence is obtained that the photolysis of acetone in the "banded" region follows the same mechanism as it does in the continuous region from about 2200 A. to 3050 A.

Pearson (60) and Glazebrook and Pearson (27) have demonstrated by use of the Paneth mirror technique that radicals are produced in the photolysis of ethyl methyl ketone. This is to be expected, since this ketone does not differ markedly from acetone. The presence of the free radicals in the photolysis of ethyl methyl ketone thus indicates that at least part of the decomposition takes place according to reaction 1, which we may write as

$$CH_3COC_2H_5 + h\nu \rightarrow CH_3CO + C_2H_5 \qquad (1a)$$

and

$$CH_3COC_2H_5 + h\nu \rightarrow C_2H_5CO + CH_3 \qquad (1b)$$

Norrish and Appleyard (52) analyzed the products of the photolysis of this ketone and found that of the hydrocarbons formed ethane, propane, and butane were present in about equal amounts, these forming 70–80 per cent of the hydrocarbon products (table 19). These authors also found 20–25 per cent of the hydrocarbons formed to be ethylene. The conditions of these studies were the use of ethyl methyl ketone vapor, above the boiling liquid, photolyzed with the full intensity of a mercury arc. Thus, the temperature of the vapor was about 80°C., at which temperature the acetyl radical is fairly stable and the propionyl radical, as will be discussed later, is quite unstable, and readily dissociates to CO and C_2H_5. Equal quantities of ethane, propane, and butane would require, from the work of Norrish and Appleyard, that 70–80 per cent of the decomposition proceed according to reactions 1a and 1b with no process of the type $CH_3COC_2H_5 \rightarrow C_3H_8 + CO$ allowable. The ethylene probably is explained by a disproportionation reaction between two ethyl radicals, possibly on the walls of the reaction cell.

TABLE 19

Photolysis of ethyl methyl ketone

Composition of products in per cent

(Norrish and Appleyard (52))

EXPT.	CO	CH_4	C_2H_4	C_2H_6	C_3H_8	C_4H_{10}	CO_2
1	45.5	5.4	9.05	13.2	27		0
2	50.0	3.05	12.3	9.8	11.6	12.5	0.55
3	51.0	4.3	4.9		39.8		

TABLE 20

The photolysis of $CH_3COC_2H_5$ *in the presence of iodine at 70°C.*

Mercury arc

Gorin (30))

PRESSURE OF $CH_3COC_2H_5$	WAVE LENGTH	PRESSURE OF I_2	RI	CO	C_3H_8
mm.	*A.*	*mm.*	*moles* × *10^6*	*moles* × *10^6*	*moles* × *10^6*
65	3130	4	22.2	1.05	0.82
70	3130	2	21.2	0.89	0.74
67	Unfiltered	4	17.2	1.24	1.12
72	Unfiltered	2	17.6	1.33	1.24

As mentioned by Steacie (77), the inefficiency of the simple bulb-to-bulb distillations used by Norrish and Appleyard probably explains the 1:1:1 ratio of the three hydrocarbons obtained rather than 1.00:1.74:0.83 for C_2H_6:C_3H_8:C_4H_{10} as calculated by Moore and Taylor (48) from statistical considerations, for equal concentrations of the methyl and ethyl radicals.

Gorin (30) has studied the photolysis of ethyl methyl ketone in the presence of iodine and arrived at essentially the same conclusions as did Norrish and Appleyard (table 20).

Since the amount of carbon monoxide formed in the presence of iodine is quite small compared to the amounts of RI, the main course of the photolysis of ethyl methyl ketone must be reactions 1a and 1b. Gorin indicates that some of the decomposition may proceed according to the one-step mechanism

$$CH_3COC_2H_5 + h\nu \rightarrow CO + C_3H_8 \qquad (44)$$

this occurring to the extent of 6.5 per cent in the unfiltered arc and 3.5 per cent at 3130 A., both figures for 70°C.

Gorin's studies of ethyl methyl ketone were not nearly as complete as his studies on acetone. Further, since he did not separate the alkyl or acyl iodides, this work should not be considered as giving the complete mechanism of the ethyl methyl ketone photolysis (table 21).

Moore and Taylor (48) have studied the proportions of the three hydrocarbons, ethane, propane, and butane, and have also compared the ratios of these with the ratios expected according to calculations on the number of collisions of the two radicals methyl and ethyl. Thus, if

TABLE 21

Comparison of the rates of photolysis of CH_3COCH_3 *and* $CH_3COC_2H_5$ *in the presence of iodine and unfiltered light*

(Gorin (30))

PRESSURE	KETONE	TIME	RI
mm.		*min.*	*moles × 10⁶*
90	$CH_3COC_2H_5$	40	24.6
90	CH_3COCH_3	40	21.7
90	$CH_3COC_2H_5$	40	22.1
90	CH_3COCH_3	40	20.6

$$CH_3 + CH_3 = C_2H_6 \qquad (a) \quad (9')$$
$$CH_3 + C_2H_5 = C_3H_8 \qquad (b) \quad (40)$$
$$C_2H_5 + C_2H_5 = C_4H_{10} \qquad (c) \quad (45)$$

the ratio of the number of collisions is

$$a{:}b{:}c = 1.00{:}1.74{:}0.83$$

The experimental values obtained were

$$C_2H_6{:}C_3H_8{:}C_4H_{10} = 1.00{:}1.54{:}1.00$$

As in the case of the high-temperature photolysis of acetone, the rate of formation of methane from ethyl methyl ketone increases markedly with temperature (table 22).

As can be seen from table 22, the amounts of propane and butane formed decrease with temperature, while ethane formation remains essentially constant. This and the methane formation are probably explained by reactions 27 and 46.

$$CH_3 + CH_3COC_2H_5 \rightarrow CH_4 + CH_3COC_2H_4 \qquad (27)$$
$$C_2H_5 + CH_3COC_2H_5 \rightarrow C_2H_6 + CH_3COC_2H_4 \qquad (46)$$
$$CH_3COC_2H_5 + h\nu \rightarrow CH_3CO + C_2H_5 \qquad (1a)$$
$$CH_3COC_2H_5 + h\nu \rightarrow C_2H_5CO + CH_3 \qquad (1b)$$

The relative importance of reactions 1a and 1b at room temperature for absorption of light near 3000 A. has been determined by Ells and Noyes (23) after comparing the quantum yields of carbon monoxide formation from acetone, ethyl methyl ketone, and diethyl ketone. From diethyl ketone only the propionyl radical can be formed. The high value of Φ_{CO} from diethyl ketone, approximately unity at 36 mm. pressure, indicates that the propionyl radical is very unstable and probably dissociates to a very large extent even after absorption of light of about 3000 A. Thus, if reaction 1b proceeded to a very large extent at room temperature the quantum yield of carbon monoxide formation would be expected to be appreciably higher for ethyl methyl ketone than for acetone. Actually, Ells and Noyes (23) found Φ_{CO} for ethyl methyl ketone to be very nearly the same as for acetone. This would indicate that reaction 1a is much more important at room temperature than is reaction 1b at the relatively long wave length of 3000 A. At lower wave lengths or at higher temperatures, reaction 1b may become as important as 1a.

TABLE 22

Percentage composition of the photolysis products of $CH_3COC_2H_5$

(Moore and Taylor (48))

TEMPERATURE, °C	90	110	155	200
H_2	0.4	0.1	0.3	0.0
CO	47.5	49.0	47.0	48.5
CH_4	3.4	4.4	9.0	18.0
C_2H_4	2.2	0.5	1.7	3.0
C_2H_6	11.5	13.0	14.5	12.5
C_3H_8	19.1	20.0	18.6	11.0
C_4H_{10}	13.2	13.0	9.0	7.0
C_4H_8	1.5		0.0	

Ells and Noyes (22) have also suggested that symmetrical ketones are formed when ethyl methyl ketone is photolyzed at about 3000 A. This would be expected on the basis of reactions 1a, 1b and reactions such as 8 and 47:

$$CH_3 + CH_3CO \rightarrow CH_3COCH_3 \tag{8}$$

$$C_2H_5 + C_2H_5CO \rightarrow C_2H_5COC_2H_5 \tag{47}$$

Reaction 47 is probably unimportant in view of the instability of the propionyl radical, but reaction 8 is quite important, as also shown in the case of the acetone photolysis.

The effect of elevated temperature on the quantum yields of carbon monoxide and methane has been studied by Ells and Noyes (22) and found not to have chain character (table 23).

At 1850–2000 A. Ells and Noyes (22) found Φ_{CO} to approach unity, as might be expected. Thus, Φ_{CO} from acetone at these wave lengths is about 0.7. With the formation of the more unstable propionyl radical, Φ_{CO} would be expected to be greater than 0.7 on the basis of the free-radical theory (see table 24).

In general, the photolysis of ethyl methyl ketone appears to follow a completely free-radical mechanism, the formation of ethylene possibly being explained by

$$2C_2H_5 \xrightarrow{\text{wall}} C_2H_4 + C_2H_6 \tag{48}$$

and at higher temperatures by reaction 27 or 46, followed by

$$CH_3COC_2H_4 \rightarrow CH_3CO + C_2H_4 \tag{49}$$

At room temperature and low pressures biacetyl is probably formed according to the reaction:

$$CH_3CO \xrightarrow{\text{wall}} \tfrac{1}{2}(CH_3CO)_2 \tag{11}$$

The presence of biacetyl in the photolysis of ethyl methyl ketone has been demonstrated by Matheson and Zabor (46), who found the green fluorescence characteristic of biacetyl when this ketone is irradiated with light from a mercury arc.

TABLE 23

Quantum yields of carbon monoxide and methane at elevated temperature

(Ells and Noyes (22))

KETONE PRESSURE	TEMPERATURE	ΦCO	$\frac{CO}{CH_4}$	ΦCH_4
mm.	°C.			
43	105–110	0.79	4.2	0.19
4.75	160–165	0.79	4.0	0.20
45	160–165	0.85	2.2	0.38
43	195–200	0.71	1.5	0.47

IV. DIETHYL KETONE

Diethyl ketone has not been studied very extensively, but the work that has been reported indicates that the photolysis of this ketone yields products that would be predicted on the basis of the free-radical theory.

Pearson and Purcell (60) have demonstrated the formation of ethyl radicals during the photolysis of diethyl ketone by use of the Paneth mirror technique. The authors isolated and identified triethylarsine dimercurichloride. Danby and Hinshelwood (17) also were able to show the presence of radicals when diethyl ketone is photolyzed by studies on the free-radical-induced polymerization of olefins (table 14 and figure 10).

Matheson and Zabor (45) have shown that the fluorescence observed when diethyl ketone is irradiated with ultraviolet light is nearly identical with that observed when propionaldehyde is irradiated. By analogy with acetone, ethyl methyl ketone, and acetaldehyde, which give similar fluorescence spectra, it seems likely that the fluorescence from diethyl ketone and propionaldehyde arises from bipropionyl, formed from the combination of two propionyl radicals.

Ells and Noyes (22) studied the photodecomposition of diethyl ketone in the regions 1850–2000 A. and near 3000 A. at 25°C. At the lower wave-length region the quantum yield of carbon monoxide formation was found to rise from 0.82 at 38 mm. ketone pressure to 0.95 at 4 mm. pressure. In view of the work of Matheson and Zabor a probable primary step in the photolysis of this ketone is:

$$C_2H_5COC_2H_5 + h\nu \rightarrow C_2H_5CO + C_2H_5 \qquad (50)$$

At about 3000 A. the quantum yield of carbon monoxide formation was found by Ells and Noyes to rise from 0.67 at 10 mm. pressure to 1.04 at 36 mm. pressure. Thus, in both wave-length regions, 1850–2000 A. and 3000 A., Φ_{CO} tends toward unity. In view of this, the propionyl radical is apparently very unstable, unless it enters into some very efficient carbon monoxide-producing reaction, and reaction 51 would appear to proceed to a large extent at all wave lengths.

$$C_2H_5CO \rightarrow C_2H_5 + CO \qquad (51)$$

The method of formation of bipropionyl has not been studied because of lack of methods for detecting very small amounts of this substance. However, at low pressures it might be expected to be formed heterogeneously at moderate intensities, as is biacetyl.

$$C_2H_5CO \xrightarrow{\text{wall}} \tfrac{1}{2}(C_2H_5CO)_2 \qquad (52)$$

The recombination of propionyl and ethyl radicals *possibly* accounts for Φ_{CO} being less than unity at higher pressures (38 mm.)

$$C_2H_5CO + C_2H_5 \rightarrow C_2H_5COC_2H_5 \qquad (47)$$

At 1850–2000 A. Ells and Noyes (22) found that 77–88 per cent of the ethyl radicals formed combined to form butane, the remainder apparently disproportionating to give ethane and ethylene, these being found in equal amounts.

$$C_2H_5 + C_2H_5 \xrightarrow[(?)]{\text{wall}} C_2H_4 + C_2H_6 \qquad (48)$$

$$C_2H_5 + C_2H_5 \rightarrow C_4H_{10} \qquad (53)$$

These equations explain very satisfactorily the observed products. An alternative reaction that might be postulated to explain the high value of Φ_{CO} is reaction 54, if this has a very low energy of activation.

$$C_2H_5 + C_2H_5CO \rightarrow C_4H_{10} + CO \qquad (54)$$

There is no advantage in the use of this equation, since if it proceeded to a large extent, ethane and ethylene would have a very small probability of formation, contrary to experiment.

A comparison of the quantum yields of carbon monoxide formation from acetone, ethyl methyl ketone, and diethyl ketone is given in table 24.

V. *n*-BUTYL METHYL KETONE

The photolysis of *n*-butyl methyl ketone was first studied by Norrish and Appleyard (52). Acetone and propylene were found to be the major products, these in approximately equivalent amounts, when the vapor over the boiling ketone (b.p. 127°C.) was irradiated with a hot mercury arc. The percentage composition of the gaseous products is given in table 25.

Experiment 3 was performed in the presence of the acetone from experiment 2, and the increase in carbon monoxide and ethane in the former is probably due to decomposition of acetone. Thus, under the conditions of the experiments made by Norrish and Appleyard, at least part of the carbon monoxide did not arise from an immediate consequence of the *n*-butyl methyl ketone.

TABLE 24

Comparison of the quantum yields of carbon monoxide formation from acetone, ethyl methyl ketone, and diethyl ketone

(Ells and Noyes (23))

WAVE-LENGTH REGION	PRESSURE	$(CH_3)_2CO$	$CH_3COC_2H_5$	$(C_2H_5)_2CO$
A.	*mm.*			
1850–2000	20	0.52	0.76	0.88
1850–2000	8	0.56	0.78	0.90
1850–2000	4	0.59	0.78	0.95
1850–2000	1	0.62	0.82	0.98*
1850–2000	0.3	0.64	0.86	0.99*
3000	10	0.095†	0.082	0.67
3000	10	0.095†	0.11	0.98
3000	21	0.095†	0.12	0.97
3000	36	0.095†	0.12	1.04

* Obtained by rough graphical extrapolation.

† This was taken as a standard after many determinations.

TABLE 25

Composition of gaseous products in per cent

(Norrish and Appleyard (52))

EXPT.	CO	H_2	CH_4	C_3H_6	C_2H_6
1	11.4	2.0	0.95	80.6	5.1
2	15.0	1.6	1.05	73.7	10.3
3	19.2	1.6	1.7	63.5	14.1

From these data Norrish and Appleyard proposed the Type II method of decomposition of this ketone

$$CH_3COC_4H_9 + h\nu \rightarrow CH_3COCH_3 + C_3H_6 \qquad \text{Type II} \quad (55)$$

Bamford and Norrish (7) found essentially the same results when *n*-butyl methyl ketone is photolyzed in paraffin solution.

Davis and Noyes (18) photolyzed *n*-butyl methyl ketone vapor at about 25°C. and 12 mm. pressure with the full mercury arc and also using approximately monochromatic 3130 A. radiation. Propylene was found to be a major product, being formed to an extent of about ten times the carbon monoxide rate of formation. Traces of ethane were found, but the methods of analysis were not accurate enough to give exact data on this compound. Although a product was found that analyzed as acetone, this substance was not found to be equivalent to propylene; rather, equal only to about 75 per cent of the latter. In view of the solubility of acetone in *n*-butyl methyl ketone it is impossible to say from this work whether acetone and propylene are really equivalent. However, since only about 1 per cent of the hexanone was decomposed, the carbon monoxide formed probably did not result from the secondary photolysis of acetone.

Wilson and Noyes (75) studied the photolysis of *n*-butyl methyl ketone at about 3130 and 2537 A. Some of their results may be summarized as follows: (*1*) at 27°C. the quantum yield of C_3-hydrocarbons (propylene plus propane) is about 0.34, independent of intensity and pressure; (*2*) the quantum yield of C_3-hydrocarbons is independent of temperature from 25–115°C. within experimental error and equal to 0.39 ± 0.02; (*3*) Φ_{C_3} is approximately the same for 3130 and 2537 A.; (*4*) the quantum yield of acetone plus other products after the removal of C_3-hydrocarbons is about 0.26 at 3130 A. and 27°C., increasing slightly with decreasing wave length; (*5*) the carbon monoxide yield is small but increases with temperature.

Wilson and Noyes also reported $\Phi_{C_2H_6}$ about 0.1, this being independent of intensity and wave length, but increasing slightly with temperature. Whether this substance, reported as ethane, is really ethane, seems quite doubtful. Norrish and Appleyard (52) and Davis and Noyes (18) found carbon monoxide in appreciable quantities, the latter using the same apparatus as employed by Wilson and Noyes. It seems probable that the −195°C. fraction obtained by Wilson and Noyes was carbon monoxide with traces of methane and ethane rather than nearly pure ethane.

Wilson and Noyes detected biacetyl in the photolysis of *n*-butyl methyl ketone. This may have resulted from the secondary decomposition of acetone, since a long run, using an unfiltered mercury arc, was made in order to obtain a measurable quantity of biacetyl. However, in view of the small amount obtained, it is also possible that the acetyl radical is formed in one of several possible minor primary processes.

$$CH_3COC_4H_9 + h\nu \rightarrow CH_3CO + C_4H_9 \qquad (56)$$

If such a process made a small contribution to the primary step, at the low pressures of ketone used by Wilson and Noyes (12 mm.), most of the acetyl radicals would be expected to form biacetyl heterogeneously according to equation 11:

$$CH_3CO \xrightarrow{\text{wall}} \tfrac{1}{2}(CH_3CO)_2 \qquad (11)$$

At higher temperatures, where the acetyl radical is quite unstable, we would expect

$$CH_3CO \rightarrow CH_3 + CO \qquad (6)$$

or

$$CH_3CO + M \rightarrow CH_3 + CO + M \qquad (7)$$

followed by

$$CH_3 \xrightarrow{\text{wall}} \tfrac{1}{2}C_2H_6 \tag{9}$$

and

$$CH_3 + C_4H_9 \rightarrow C_5H_{12} \tag{57}$$

Whether ethane is formed in more than negligible amounts has still to be proved; hence the value of reaction 56 in explaining the formation of biacetyl or other products is to be questioned. Wilson and Noyes have used reaction 56 to explain the products of the photolysis of butyl methyl ketone, especially the inequality of propylene and acetone. These authors proposed that equation 56 also be followed by

$$C_4H_9 \rightarrow C_3H_6 + CH_3 \tag{58}$$

to explain the formation of propylene. Further, these authors suggested

$$CH_3COC_4H_9 + h\nu \rightarrow CH_3COCH_2 + CH_3CH_2CH_2 \tag{59}$$

followed by

$$CH_3COCH_2 + CH_3 \rightarrow CH_3COCH_2CH_3 \tag{60}$$

$$2CH_3CH_2CH_2 \rightarrow C_3H_6 + C_3H_8 \tag{61}$$

However, Moore and Taylor (48), studying the mercury-photosensitized hydrogenation of propylene, found reaction 61 to go to an extent of only about 26 per cent compared with 64 per cent of reaction 62:

$$2CH_3CH_2CH_2 \rightarrow C_6H_{14} \tag{62}$$

Moore and Taylor (48), studying the mercury-photosensitized hydrogenation of butene, found only 1 per cent propane, 29 per cent butane, and 66 per cent octane, indicating that the extent of reaction 58 is small, at least at low temperatures, compared with reactions 63 and 64.

$$2CH_3CH_2CH_2CH_2 \rightarrow C_4H_{10} + C_4H_8 \tag{63}$$

$$2CH_3CH_2CH_2CH_2 \rightarrow C_8H_{18} \tag{64}$$

Recent work by Davis and Noyes (19) indicates that at low temperatures the main reaction must be that given by reaction 55; i.e., a direct split into the final products. Up to 100°C. acetone and propylene were found to be formed in equal amounts. From this work $\Phi_{C_3H_6} = 0.45 \pm 0.05$ at 100°C., as compared with a value of 0.39 ± 0.02 found by Wilson and Noyes (75). Studies on the effect of temperature (19) showed that acetone was formed up to 300°C. and therefore could not be formed by the reaction 8, owing to the great instability of the acetyl radical.

$$CH_3CO + CH_3 \rightarrow CH_3COCH_3 \tag{8}$$

At higher temperatures, above 100°C., the radical forming primary process 56 appears to become increasingly important, the yields of carbon monoxide, methane, ethane, propylene, and C_4-hydrocarbons increasing considerably. At these higher temperatures reaction 58 becomes quite important.

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THE SANDMEYER REACTION

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I. INTRODUCTION

During an attempt in 1884 to prepare phenylacetylene from benzenediazonium chloride and cuprous acetylide, Sandmeyer (53) obtained chlorobenzene in good yield, and he subsequently established that the active agent in promoting the reaction was cuprous chloride. Sandmeyer explained his result by two equations, *viz.:*

$$C_6H_5N{=}NCl + Cu_2Cl_2 = C_6H_5N(CuCl)N(CuCl)Cl$$

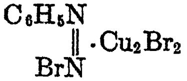
and

$$C_6H_5N(CuCl)N(CuCl)Cl = N_2 + Cu_2Cl_2 + C_6H_5Cl$$

He suggested that the intermediate compound with cuprous chloride might be the ephemeral yellow precipitate obtained when the diazobenzene chloride was added to a solution of cuprous chloride in hydrochloric acid; he stated also that cupric and ferrous chlorides did not act in this way. In analogous fashion, cuprous bromide was found to cause the replacement of the diazo group by bromine in benzenediazonium bromide (54), while cuprous cyanide similarly effected the substitution of the diazo by the cyano group (54). This facilitation by cuprous salts of replacement of the diazo group by halogens or other groups was thought by Sandmeyer to be specific. This viewpoint was held by chemists for well over 50 years and still has its adherents (71); the reaction in consequence became universally known as *The Sandmeyer Reaction*. It is of great interest, however, that although the diazo compounds were decomposed when heated with hydrochloric or hydrobromic acid, with evolution of nitrogen, the yields of halogeno products were poor (nevertheless hydriodic acid appeared to convert aryldiazonium salts to the corresponding aryl iodides even in the cold (19)), and that the replacement of the diazo group by fluorine was either uninfluenced by copper, or, if influenced, the effect may have been unfavorable, since Hantzsch and Vock (28) obtained *p*-bromophenol by acting on *p*-bromobenzenediazonium fluoride with copper. As already noted, Sandmeyer (53) himself tried cupric and iron salts in place of cuprous salts and found them to be without action, while Blumberger (2) reported that cupric salts merely accelerated decomposition. The first real challenge to cuprous chloride or bromide, so far as replacement of the diazo group by chlorine or bromine was concerned, came in 1923, when Contardi and his collaborators (9) found that the decomposition of substituted naphthalenediazonium salts (principally from nitronaphthylamines) was effected by cupric salts, sometimes with better yields than when cuprous salts were used.

II. VIEWS ON THE MECHANISM OF THE SANDMEYER REACTION PRIOR TO 1941

It has already been noted that Sandmeyer (53) attributed the efficacy of cuprous halides or cyanides to their formation of an unstable intermediate compound with the diazonium salts, and although he himself was unable to isolate any such intermediate product (54), yet Hantzsch (25) prepared what he termed the *syn*-diazobenzene cuprobromide,

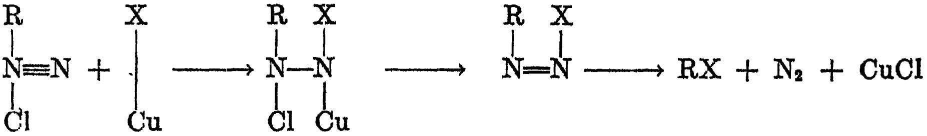


$$\begin{array}{l} C_6H_5N \\ \quad \| \cdot Cu_2Br_2 \\ BrN \end{array}$$

which was obtained in large, reddish yellow needles and decomposed on addition of water into cuprous bromide, nitrogen, and bromobenzene, no diazonium bromide being re-formed; *p*-bromo-*syn*-diazobenzene cuprobromide, $BrC_6H_4N{=}NBr \cdot Cu_2Br_2$, was similar to the foregoing but of a dark red color. Further, Lellmann and Remy (49) were also successful in isolating a red crystalline substance from β-naphthalenediazonium bromide and cuprous bromide, and these double salts all differed from colorless diazonium salts not only in being strongly colored (in contrast likewise to cuprous halides, which are colorless), but also in their stability and ready mode of decomposition to give the aryl or naphthyl halides. In Sidgwick's *Organic Chemistry of Nitrogen* (62), the viewpoint is expressed that in the double salts the diazo nitrogen is probably coördinate to the copper, a union which seems likely to alter the type of decomposition.

The halogen atom which replaces the diazo group was held by Hantzsch and Blagden (27) to be derived from the cuprous halide, on the evidence that *p*-chlorobromobenzene was obtained from *p*-bromobenzenediazonium bromide and cuprous chloride, while conversely *p*-dibromobenzene was obtained from *p*-bromobenzenediazonium chloride and cuprous bromide in methyl sulfide solution. Incidentally, this evidence received some support from experiments by Hodgson and Walker (46), who found that the replacement of the diazo group by chlorine or bromine could be carried out even in concentrated sulfuric acid solution, although in these cases the corresponding halogen acid was also present. Saunders (58) remarks, however, that it is well known that in replacing the diazo group by bromine a higher yield is often obtained if diazotization is carried out with hydrobromic acid than if hydrochloric acid is used; this experience will be dealt with later (Section IV, B).

To explain the above results, Hantzsch (24) proposed a general formulation based on the well-known formation of colored intermediate addition compounds, and analogous to the colored solid diazo halides, which comprehended also the preparation of the nitriles and other analogous derivatives, *viz.:*



$$\underset{\mathrm{Cl}}{\overset{\mathrm{R}}{\mathrm{N}}}{\equiv}\mathrm{N} + \underset{\mathrm{Cu}}{\overset{\mathrm{X}}{|}} \longrightarrow \underset{\mathrm{Cl}}{\overset{\mathrm{R}}{\mathrm{N}}}-\underset{\mathrm{Cu}}{\overset{\mathrm{X}}{\mathrm{N}}} \longrightarrow \overset{\mathrm{R}}{\mathrm{N}}=\overset{\mathrm{X}}{\mathrm{N}} \longrightarrow \mathrm{RX} + \mathrm{N_2} + \mathrm{CuCl}$$

Addition complex *syn*-Diazo compound

R = aryl radical; X = halogen, cyano, cyanate, etc.

Groggins (20) has stated that copper is not the only metal that can advantageously be employed for replacement of the diazo group by halogen, and that when precaution is taken to ensure a maximum concentration of chloride ions, saturated solutions of calcium chloride, zinc chloride, and zinc-ammonio chloride often give equally good results. In the corresponding decomposition to form nitriles, when a diazo solution is added to one of cuprous cyanide in potassium cyanide, a complex containing copper separates which decomposes on warming, and, while the yields are usually quite good, nevertheless they are improved by substituting nickel cyanide, $Ni(CN)_2$, for cuprous cyanide; this reaction is apparently specific for copper and nickel, since the cyanides of iron, chromium, zinc, and other metals are stated to be quite useless (48). A good recipe for the preparation of cuprous cyanide for use in the Sandmeyer reaction is given by Barber (1), who employs the reaction

$$2CuSO_4 + 2NaCN + NaHSO_3 + H_2O \longrightarrow 2CuCN + 3NaHSO_4$$

and who states that the cuprous cyanide thus prepared can be kept indefinitely.

Groggins (20) envisaged the mechanism of the Sandmeyer reaction as follows: The aryldiazonium compound first forms an addition compound with cuprous salts, resulting in a complex cation which subsequently undergoes decomposition. In this process, it is not the halogen inherent in the diazonium compound or the halogen ions of the reaction solution, but the halogen of the copper halide by virtue of its closer proximity to the aryl radical that enters the nucleus. That quantitative yields are never obtained in carrying out this reaction is explained by Groggins on the hypothesis that some of the diazonium salt reacts with hydroxyl ions before the addition complex is formed.

The kinetics of the Sandmeyer reaction were first investigated by Waentig and Thomas (68), who found that the reaction was not monomolecular, if the assumption be made that the measurable reaction is the decomposition of the cuprous diazo double compound, but that the acid concentration has considerable influence on the decomposition velocity—in direct contrast to the simple formation of phenols from diazonium compounds. Sandmeyer (56) had further supported his hypothesis of a double compound by the formation of nitrobenzene from aniline, and stronger evidence was obtained by Hantzsch (27), who found that the double salt of benzenediazonium nitrate and mercury nitrite, $(C_6H_5N_2NO_3)_2 \cdot Hg(NO_2)_2$, afforded nitrobenzene almost entirely when decomposed.

More recently Eichenberg (12) pointed out that two factors of decisive importance had to be considered: (*1*) the activation energy required to discharge the chloride ions; (*2*) the operation of the law of mass action. It is evident, he states, that in a solution containing hydroxyl as well as halogen ions, the former will move with a greater velocity and react more readily, and therefore play a prominent rôle in the formation of phenols; on the other hand, from the law of mass action, it would be expected that in a solution containing different ions, each will react in proportion to its concentration. From the consideration of molecular or ionic concentrations alone, practically theoretical yields of aryl halide would be expected when maximum concentrations of hydrochloric acid and zinc chloride were used in the Sandmeyer reaction, but this does not occur, and a discussion of this anomaly is given later in the paper.

Gattermann (17) believed free copper to be as effective as the cuprous salts, but Saunders (58) pointed out that when copper is used to bring about halogen replacement it is the cuprous ions first formed by the action of the acid on the copper which really do the work, and that the reaction only proceeds in acid solution and hence appears to be a reaction of the diazonium salt. On the other hand, Hantzsch considered that any intermediate stages were derived from the *syn*-diazo hydroxide.

This paper will deal only tentatively with the Gattermann reaction, which is still being experimentally studied by the author, but it may be pointed out that *o*- and *p*-dinitrobenzenes are obtained in good yield by the decomposition of *o*- and *p*-nitrobenzene fluoroborates in aqueous solution containing sodium nitrite by means of copper powder and in the absence of acid (63). Naturally, however, in the presence of hydrochloric or any other acid, cuprous salts will be formed incipiently as envisaged by Saunders, so that Hantzsch's viewpoint of a double rôle would appear to have considerable justification.

According to Erdmann (14), the normal decomposition of the diazo cuprous chloride compound takes place rapidly and smoothly only above a certain temperature, which is different for each compound; these temperatures are about 0°, 27°, and 30–40°C. in the case of cuprous chloride derivatives of diazobenzene, *o*-diazotoluene, and *p*-diazotoluene, respectively. Below these points the evolution of nitrogen takes place too slowly and is incomplete, part of the diazo cuprous chloride compound being converted to an azo compound by the liberated cuprous chloride. It has been found that the quantity of cuprous chloride required may be reduced to 1/21 and 1/28 mole per mole of amine without materially reducing the yield of chlorobenzene (66.2–63.5 per cent) and *m*-chloronitrobenzene (78.5–73.2 per cent), respectively (66). This is considered to be due to the diazobenzene chloride becoming first reduced to phenylhydrazine by the cuprous chloride, which subsequently becomes cupric chloride; the phenylhydrazine is then oxidized in the presence of hydrochloric acid to chlorobenzene by the cupric chloride, and the cuprous chloride would then be reformed to play the same part again. In an assumed confirmation of this explanation, it is found that phenylhydrazine is oxidized to chlorobenzene under the conditions named by both cupric and ferric chlorides, but the fact remains that no phenylhydrazine can be detected in the Sandmeyer reaction, and this is ascribed to its momentary existence. Against this viewpoint, it is worthy of emphasis that a copper salt is not necessary in the preparation of iodobenzene by this method, and this is explained by the fact that hydriodic acid is itself a reducing agent (69). Later it will be shown that iodobenzene is not formed in the presence of reducing agents.

III. INDUSTRIAL APPLICATIONS OF THE SANDMEYER REACTION

The Sandmeyer reaction has not as yet received many technical applications, but a few may be cited: (*1*) the preparation of 2,6-dichlorotoluene from 2,6-chlorotoluidine and of 2-chlorobenzaldehyde (15); (*2*) the production of 1-chloro-8-naphthol-3,6-disulfonic acid from H acid (5); and (*3*) the preparation of halogenoanthraquinones (21). An electrolytic process (67) has also been based on the Sandmeyer method, and has been devised to avoid the preparation and recovery of cuprous salts in large-scale technical operations. It consists in electrolyzing a solution of a diazo compound to which cupric chloride has been added, where a thick copper wire is used as the anode and a cylinder of sheet copper as the cathode. With a current density of 2.1 amperes per square decimeter, and an E. M. F. of 10 volts, nitrogen was evolved and a yield of 64 per cent of the theoretical yield of chlorobenzene was obtained. Apparently, however, such processes have not enjoyed wide technical use.

IV. AN INTERPRETATION OF THE SANDMEYER REACTION

A. Foreword

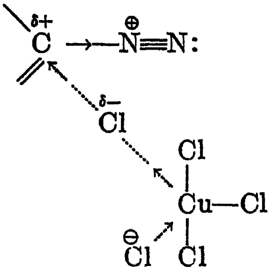
Although various suggestions have been noted in the introduction for the mechanism of the Sandmeyer reaction, all of them are invalid for one reason or another. It is stated by Hammett (23) that "*the characteristic effect of cuprous compounds in the Sandmeyer reaction has no obvious explanation,*" and by Sidgwick (61) that the action of metallic copper is not understood. In an attempt to solve this problem, Hodgson, Birtwell, and Walker (34) have submitted an interpretation of the Sandmeyer reaction and its variations, based on a mechanism whereby the exceedingly stable halogen ions involved are rendered reactive (i.e., are somewhat destabilized) by cuprous salts acting purely as catalysts, which convert them into nucleophilic (anionoid) reagents by a temporary sharing of their stable octets. This interpretation was then found to explain: *(1)* the similar catalytic action of cupric, iron, and cobalt salts; *(2)* the experimental results of Hantzsch and Blagden (27); *(3)* the anomalous behavior of fluorine compounds; *(4)* the ease of reaction between diazonium compounds and iodides in the absence of copper salts; and *(5)* the action of halogen acids on diazonium compounds. Also, the formation of diaryls (16, 57, 64) in the Sandmeyer reaction was shown to be a consequence of an oxidation–reduction mechanism involving conversion of cuprous into cupric salts.

B. Mixed Sandmeyer reactions

Eight arylamines (2-, 3-, and 4-nitroanilines, *p*-chloro- and *p*-bromoanilines, anthranilic acid, *p*-phenylenediamine, and benzidine) were each diazotized by the glacial acetic acid–nitrosylsulfuric acid procedure of Hodgson and Walker (46), and the solutions of the diazonium salts thus obtained were decomposed in two ways: (*a*) by cuprous chloride dissolved in concentrated hydrobromic acid and (*b*) by cuprous bromide dissolved in concentrated hydrochloric acid. Taking the example of *p*-nitroaniline as typical of the eight, it was found for (*a*) that the weights of 4-chloronitrobenzene and 4-bromonitrobenzene which resulted were in the ratio of *ca.* 4:96, while for (*b*) the ratio had shifted to 60:40. This distribution superficially contradicts the results of Hantzsch and Blagden (27), which indicated the halogen to come from the cuprous salt, since in (*a*) the halogen apparently comes directly from the acid. It will be shown below, however, that there is complete harmony between all the data.

To interpret these results, the dissolution of cuprous chloride and bromide in hydrobromic and hydrochloric acid, respectively, is assumed to form complex anions, *viz.*, $[CuClBr_3]^{---}$ and $[CuCl_3Br]^{---}$, and in these complexes the bromine (see later) will be more reactive than the chlorine. The reason (35) for this formulation of the anions is that the older formula for cuprous chloride, Cu_2Cl_2, has now been replaced by CuCl, since there is little reliable evidence for the Cu—Cu link in cuprous compounds (because in these compounds where detailed structure is known, the cuprous atoms are too far apart for such direct linkage); moreover, such structures usually enable the cuprous atoms to acquire seven electrons and thus attain the krypton structure (10, 18, 50, 72). The active complex cuprous anion in a solution of cuprous chloride in hydrochloric acid should therefore be represented by $[CuCl_4]^{---}$. This formulation will also be in harmony with the established compound, $K_3Cu(CN)_4$, which Grossmann and van der Forst (22) have shown to represent the most stable type of complex cuprous cyanide.

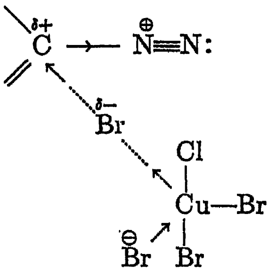
The mechanism of decomposition of aryldiazonium chlorides by cuprous chloride dissolved in hydrochloric acid could then be envisaged as follows: Attraction of the complex anion $[CuCl_4]^{---}$ to the diazonium cation, and then



A

decomposition of the complex A, in the sequence release of an electron at the carbon with separation of neutral chlorine, evolution of nitrogen and linkage of neutral chlorine to the aryl radical, followed by the rehabilitation of the complex anion by coördination with an external chloride anion.

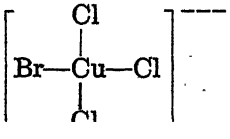
For decomposition (*a*) above, the mixed complex B will react by a bromine atom at the cationoid carbon of the diazonium ion, since bromine is more reactive



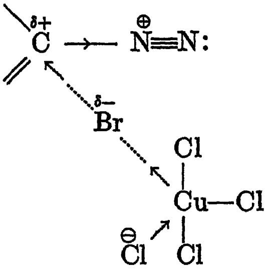
B

than chlorine, owing to its exterior electrons being less under control by the positive nucleus (*cf.* also Schwechten (60)); moreover, release of a bromide ion from the copper to the cationoid carbon entails immediate replenishment of bromine to the copper from the bromide ions in solution, so that a preponderance of bromo compound will be formed (*ca.* 96 per cent in the case of *p*-nitroaniline). It is reasonable to assume, however, that some of the complex ions will be in an environment where the chlorine can be more reactive than the bromine, and in this case a chloro compound will result (*ca.* 4 per cent for *p*-nitroaniline).

For decomposition (*b*) above, the complex anion



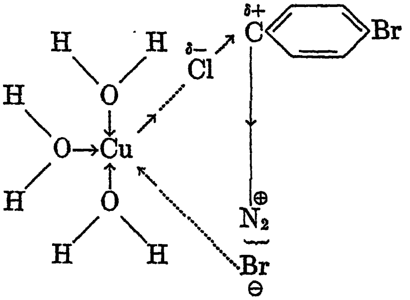
will react by a bromine atom at the cationoid carbon of the diazonium ion, and

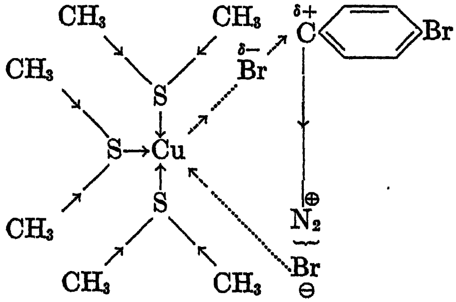


the diazo nitrogen will be displaced initially by bromine, but in this case the depleted copper anions will be replenished by chloride anions from the solution, whereby chloro complexes will now gradually take the place of the cuprous chlorobromo anions and the resulting mixture of chloro- and bromoaryls will in consequence possess a more even distribution of the halogens than in (*a*). Actually, the chloro compound predominates (60–70 per cent), as would be expected.

C. The Hantzsch–Blagden data (27)

The experiments of Hantzsch and Blagden to prove that the halogen came from the complex were as follows: The diazonium salt (1 mole) was treated with cuprous halide (1 mole) in aqueous suspension and also in methyl sulfide solution. When (*a*) cuprous chloride reacted with *p*-bromobenzenediazonium bromide, the main product was *p*-chlorobromobenzene, and (*b*) when cuprous bromide reacted with *p*-bromobenzenediazonium chloride, the main product was *p*-dibromobenzene. From the present author's point of view (see above), it would appear that an incipient complex anion is formed in each case between the cuprous halide and coördinating solvent ions, *viz.*, water and methyl sulfide, respectively. Reaction would then occur *via* (*a*) chlorine and (*b*) bromine, with formation of *p*-chlorobromobenzene in (*a*) and *p*-dibromobenzene in (*b*).





The halogen ions from the diazonium salt will then replace the departing halogen of the complex anion, and so tend to form the other halogen derivative in minor yield, as found by Hantzsch and Blagden. The contention of these authors that the halogen came from the cuprous salt is therefore valid, but if they had carried out the mixed Sandmeyer experiments described by the present authors (35), especially where nearly 100 per cent of bromo compound resulted when cuprous chloride was employed in hydrobromic acid solution, an erroneous deduction that the halogen came directly from the acid would have been well nigh irresistible. The more vigorous reaction obtained when methyl sulfide was employed in homogeneous medium would be expected from the greater abundance of the copper complex and from its greater negativity, which is due to the more permanent coördination with methyl sulfide than with the polarized water molecules. A series of experiments (34) with *p*-nitrobenzenediazonium chloride and a constant weight of cuprous bromide dissolved in various amounts of concentrated hydrochloric acid showed that as the cuprous bromide concentration decreased the amount of replacement by chlorine increased, though not proportionately, a result which is in accordance with the hypothesis now advanced to interpret the Sandmeyer reaction.

D. Copper as catalyst in the Sandmeyer reactions

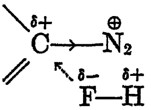
Gattermann's substitution of metallic copper for cuprous halide (17) was tested (34) for mixed halogen acids by: (*a*) diazotizing *p*-nitroaniline in hydrochloric acid, adding the equivalent of hydrobromic acid, and treating the mixture with copper powder; (*b*) diazotizing *p*-nitroaniline in concentrated sulfuric acid, adding concentrated aqueous solutions of sodium chloride and sodium bromide, and then decomposing the mixture with copper powder. For comparison, *p*-nitroaniline was diazotized by the Hodgson–Walker method (46), and the product treated separately with a solution of (*c*) cuprous bromide and (*d*) cuprous chloride in a mixture of hydrochloric and hydrobromic acids. In both (*a*) and (*b*), 4-bromonitrobenzene was formed in over 85 per cent yield together with *p*-nitrophenol. Almost theoretical yields of halogenonitrobenzenes were obtained in (*c*) and (*d*), the distribution being (*c*) 96 per cent of 4-bromo- and 4 per cent of 4-chloronitrobenzene, (*d*) 94 per cent of 4-bromo- and 6 per cent of 4-chlorobenzene; *p*-nitrophenol was absent in both experiments. The overwhelming replacement of the diazo group by bromine was thus in accord with previous experience. The effect of the water content in (*a*) and (*b*) is shown by the formation of *p*-nitrophenol.

E. Mixed Sandmeyer reactions with cuprous iodide

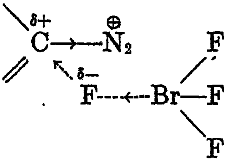
When solutions of *p*-nitroaniline diazotized by the Hodgson–Walker method (46) were poured into suspensions of cuprous iodide in hydrochloric acid and in hydrobromic acid, the yields of 4-iodonitrobenzene were 80 per cent and 75 per cent, respectively. It follows that in the incipiently formed cuprous chloroiodo and cuprous bromoiodo complexes, the iodine is much more reactive than bromine or chlorine, and the bromine is more reactive than the chlorine, since the yield of 4-bromonitrobenzene exceeded that of 4-chloronitrobenzene under equivalent conditions. These experiments are also in accord with those of Hantzsch and Blagden (27), who reported that cuprous iodide furnished iodo derivatives only with various diazo chlorides and bromides.

F. The anomalous case of hydrofluoric acid

In aqueous solution hydrofluoric acid is regarded as being present mainly as molecules of H_2F_2 (largely ionized as $H^+\{HF_2^-)$ with only a small amount of HF (probably non-ionized). These properties account for the ineffectiveness of copper as a catalyst in the ordinary Sandmeyer or Gattermann reactions, since the affinity of fluorine for hydrogen and the stability of the $\overset{\ominus}{HF_2}$ anion apparently prevent cuprous ion formation. The small yield of fluoro compound (about 25 per cent from diazotized *m*-nitroaniline (42)) would therefore appear



to be due to attack of cationoid carbon by non-ionized HF, the $\overset{\ominus}{HF_2}$ anion being unreactive, like the chloride and bromide anions. When, however, a dry bromofluoride is decomposed by heating, e.g., $R\overset{\oplus}{N_2}\}Br\overset{\ominus}{F_4}$. the same kind of process occurs as in the cuprous salt mechanism, *viz.*, a polarized fluorine atom in the



$Br\overset{\ominus}{F_4}$ ion attacks the cationoid carbon, like chlorine in the $[CuCl_4]^{---}$ complex, with resulting evolution of nitrogen and BrF_3, and subsequent replacement of the diazo group by fluorine. Such a reaction is exactly analogous to the heating of dry diazo perbromides, which affords improved yields of bromo compounds over those obtained in aqueous solution (58, page 154).

The poor yields of halogeno compounds obtained when diazonium salts are heated with even the most concentrated hydrochloric and hydrobromic acids would, by analogy with hydrofluoric acid, be due to the simultaneous reaction of water (fast) and non-ionized acid (slow, since the concentration is small). Blumberger's observation (3) that the addition of chlorides to decomposing dilute hydrochloric acid solutions of diazobenzene chloride somewhat increased the yield of chlorobenzene gives supporting evidence of the previous point of view, since the concentration of the non-ionized but very unstable hydrogen chloride would be slightly increased. The enormous excess of chloride anions present would indicate direct attack by them to be improbable, and, in the case of very concentrated hydrobromic acid, the bromide ions would exceed the amount of water present.

G. The influence of chloride ion on the decomposition of diazonium salts

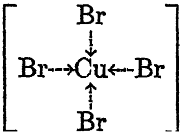
To test the conclusion above that the chloride ion plays little or no part in the decompositions of diazonium salts, Hodgson and Sibbald (44) examined the influence of chloride ions on the decomposition of neutral *p*-nitrobenzenediazonium chloride and of *p*-nitrobenzenediazonium cobaltinitrite. Most of the decompositions of diazonium chlorides hitherto reported have been studied in acid solutions (2, 3, 34, 36, 43), and have been found to vary greatly in the presence of hydrochloric acid either with sodium chloride or with the addition of various catalysts: in particular, when the acid concentration was altered from 29 to 6 per cent, the yield of *p*-chlorobenzene was altered from 56 to 17 per cent. Sodium chloride was present in all these experiments, and in the case of the 56 per cent yield the solution was saturated with it, so that it appeared of interest to ascertain whether metallic chlorides, i.e., the chloride anion, took part in the decomposition, or whether the *p*-chloronitrobenzene formed was mainly due to the unionized hydrogen chloride (36). For this purpose, the acid solution of *p*-nitrobenzenediazonium chloride was exactly neutralized with the appropriate base, and the corresponding metallic chlorides added in equivalent amounts to secure exactly parallel sets of decompositions. Except in the cases of calcium and zinc chlorides, it was not possible to ensure complete neutrality throughout the decomposition on account of the hydrochloric acid formed by the decomposing diazonium chlorides, but in these cases, excess of calcium carbonate and of zinc oxide, respectively, was added to maintain neutrality. Comparison of the data obtained, with and without filtering off the excess base, showed that the acid formed was without significant effect upon the yield, and consequently was ignored in the other experiments where its presence was unavoidable. It was found that in neutral decompositions, the addition of relatively large amounts of metallic chlorides caused no significant rise in the yield of chlor compound above that obtained when no salts were added. There was, howeve one exception, *viz.*, that when cupric chloride was added to the solution of diaz compound neutralized with calcium carbonate the yield rose to 85 per cen This remarkable result was obtained with about one-third of the equivalen amount of chlorine present in solution, which would appear to indicate the presence of a complex hydrated cupric chloride anion in the experiments in acid solution (this will be discussed later (43)). The results indicate quite clearly that the chloride ion plays little if any part in the substitution of chlorine for the diazo group, since the only cases of appreciable substitution are those in which there is the possibility of the existence of covalent (or partly polarized) chlorine, *viz.*, with hydrogen, cupric, and calcium chlorides. The behavior of zinc chloride, which unexpectedly did not follow that of calcium chloride, was probably due to the same cause as that which differentiates it from that of cuprous or cupric salts in the decomposition of its complex salts with diazonium compounds, *viz.*, that bound, anionoid, and therefore reactive water is further from the zinc than the polarized chlorine atoms (i.e., the reverse of the case of crystalline $K_2CuCl_4\cdot 2H_2O$, which will be discussed later). Support for this view is to be found in Hodgson and Foster's observation (38) that decomposition of the zinc diazonium complexes in phenol results in *ca.* 50 per cent chlorine substitution.

p-Nitrobenzenediazonium cobaltinitrite has been shown by Hodgson and

Marsden (41) to undergo replacement of the diazo by the nitro group when treated with sodium nitrite, copper sulfate, and cuprous oxide. When the decomposition was carried out in the presence of sodium chloride and copper sulfate in neutral solution, the product formed was pure *p*-chloronitrobenzene, showing that the cupric chloride has been more effective than the cobaltinitrite, or alternatively that polarized chlorine is more reactive than the polarized nitro group in the cobalt complex. When the cobaltinitrite was decomposed by sodium nitrite and sodium chloride in the absence of copper, the product was entirely *p*-dinitrobenzene, showing that the fully ionized chlorine was without effect on the decomposition, which took place *via* the polarized nitro group in the complex (see below) and was in accord with the results obtained with metallic chlorides mentioned above. Decompositions with sodium nitrite and sodium chloride in conjunction with copper sulfate gave mixtures of *p*-chloronitro- and *p*-dinitrobenzenes.

H. The action of cupric salts on diazonium compounds

Sandmeyer (53) has reported that cupric salts are without the same action as cuprous salts, and Blumberger (2) that cupric salts merely accelerate decomposition. These statements have long been true only in degree, since good yields of bromo compounds are formed when amines are diazotized in sulfuric acid solution and then treated with a solution of sodium bromide (no action alone) and copper sulfate. Further, Hodgson (32) found that when 3-bromo-4,6-dinitroaniline was diazotized in sulfuric acid solution and then treated with uprous chloride and hydrochloric acid, 1,6-dichloro-3-bromo-4-nitrobenzene was found in the cold and 1,4,6-trichloro-3-bromobenzene in the hot reaction. However, treatment at or below 0°C. with sodium chloride (no action alone), ollowed by pure copper sulfate, gave the required 1-chloro-3-bromo-4,6-dinitrobenzene in almost quantitative yield. From these results it would appear to follow that the Sandmeyer reaction is dependent on the degree of positivity of the diazonium ion, and that when this positivity is superlatively intense, as in the 3-bromo-4,6-dinitrobenzene-1-diazonium ion, cupric salts are able to bring about replacement by chlorine. Further, cupric salts have been shown (9) to react readily like cuprous salts with nitronaphthalenediazonium ions, and thereby provide supporting evidence for the point of view above with respect to positivity and facility of replacement. Germane also is Saunders' opinion (58) that the Sandmeyer reaction is a characteristic of the diazonium salt, based on his viewpoint at the time that it proceeded only in acid solution. Since then, the experiments described in the section above have shown the Sandmeyer reaction to take place with cupric chloride in neutral solution.



$$\left[\begin{array}{c} \mathrm{Br} \\ \downarrow \\ \mathrm{Br}\rightarrow\mathrm{Cu}\leftarrow\mathrm{Br} \\ \uparrow \\ \mathrm{Br} \end{array}\right]$$

The mechanism of the catalytic action of cupric salts can be assumed to differ in no wise from that of cuprous salts, *viz.*, the formation of complex cupric anions, which, while of ephemeral existence, nevertheless serve to activate the otherwise stable halogen ions.

I. The comparative behavior of cuprous, cupric, and ferric chlorides as catalysts in strong acid solution

Experiments have been described by Hodgson, Birtwell, and Walker (36), in which dilute sulfuric acid solutions of *p*-nitrobenzenediazonium sulfate were decomposed in the presence of ferric chloride alone, and also with copper sulfate and sodium chloride, which gave results indicating clearly the catalytic influence of the copper salt and of the chloride-ion concentration in promoting the formation of the complex anion. In the decompositions of *p*-nitrobenzenediazonium chloride, a 41 per cent yield of 4-chloronitrobenzene was obtained in fairly concentrated hydrochloric acid alone, and, under more concentrated conditions, the yield increased to 54.4 per cent (37). Since increase in hydrochloric acid concentration resulted in increased replacement by chlorine, it follows that this must be ascribed to the increase in concentration of the anionoid complex in the stronger acid, the simultaneous reaction by anionoid water with production of *p*-nitrophenol showing a corresponding decrease. This result is in substantial agreement with the earlier views of Hantzsch (26) and the deduction of Waters (70) on the activity of unionized hydrogen chloride in the decomposition of benzenediazonium chloride in very concentrated aqueous solution, or in concentrated hydrochloric acid, where the yield of chlorobenzene falls off very rapidly as the solution is diluted. The experiments with diazonium chlorides in concentrated acid were therefore carried out with a series of metallic chlorides whose metals are prone to form complex halogeno anions, *viz.*, those of aluminum, antimony, chromium, cobalt, copper (cuprous and cupric), iron (ferric), mercury (mercuric), nickel, tin (stannic), and zinc. Since sodium chloride under the concentrated conditions was also present in the solid form by precipitation during diazotization with sodium nitrite, all the solutions were saturated thereby, and all the decompositions were thus carried out under practically the same conditions, except for the small differences in boiling points of the various solutions.

A noteworthy feature of the decomposition of diazotized *p*-nitroaniline in strong acid solution is the efficiency of ferric and cupric chlorides as catalysts, a 77.6 per cent yield of *p*-chloronitrobenzene being obtained, i.e., the same as that with cuprous chloride under like conditions. On the other hand, aluminum chloride only slightly increased the yield over the blank, and calcium, chromium, mercuric, and zinc chlorides had no effect upon it. In marked contrast, antimony trichloride reduced the yield to 37.1 per cent (blank, 54.4 per cent), a similar effect having already been observed under more dilute conditions (41 per cent to 32 per cent), and it would appear that antimony chloride activated anionoid water preferentially to chlorine.

The experiments with cobalt chloride had an interesting bearing on its ionic state. When anhydrous, or in solutions of high chloride-ion concentration, it is blue, owing, it is generally supposed, to the $[CoCl_4]^{--}$ anion; such solutions afforded a 70.9 per cent yield of chloro compound, whereas the more aqueous pink solution, containing $Co(H_2O)_6Cl_2$, gave a yield of only about 10 per cent, i.e., the same as that in hydrochloric acid at the same dilution with no cobalt present. Also, when *p*-nitroaniline was diazotized by the Hodgson and Walker method (46) in concentrated sulfuric–glacial acetic acid solution, and then decomposed by the blue solution of cobalt chloride in concentrated hydrochloric acid, the yield was decreased to that of the blank (40.5 per cent), showing that the dissociation of the complex ion in the composite solution had inhibited its efficiency for replacement of the diazonium group by chlorine. These experiments with cobalt chloride support the interpretation given for the previous results with cupric chloride, and with copper sulfate and sodium bromide (36), that the decomposition is due to the complex $[CuCl_4]^{--}$ or $[CuBr_4]^{--}$ anion. In this connection, it is of interest that anhydrous cupric bromide is black, this color being ascribed to $Cu[CuBr_4]$, and that the addition of sodium bromide to copper sulfate produced a color change from blue to green-brown, due probably to the formation of the $[CuBr_4]^{--}$ anion; in like manner, cupric chloride is held be $Cu[CuCl_4]$ (13). The great efficiency of ferric chloride in concentrated acid solution is also due to the formation of a complex anion, and the older viewpoint that there is considerable evidence for the existence of $HFeCl_4$ has received support from evidence discussed in a later section.

Similar experiments with *m*-nitroaniline (37) usually gave somewhat lower yields of *m*-chloronitrobenzene, but in every case in the same order of efficiency as those with *p*-nitroaniline (37). In these reactions also, ferric and cupric chlorides proved to be as efficient as cuprous chloride, and the decreased yields were to be anticipated in view of the lower positivity of the significant carbon atom to which the diazonium group is attached.

J. *The comparative behavior of cuprous, cupric, and ferric chlorides as catalysts over a wide range of hydrochloric acid concentrations*

Experiments have been carried out by Hodgson and Sibbald (43) in which *p*-nitrobenzenediazonium chloride was decomposed over a range of hydrochloric acid concentrations with cuprous, cupric, and ferric chlorides as catalysts, and with cuprous and cupric chlorides together. With ferric chloride, an almost proportional decrease in yield of *p*-chloronitrobenzene occurs on dilution.

When the percentage yield according to the amount of water present is plotted, the plot for cuprous chloride has a maximum at *ca.* 18.6 per cent of acid, while that for cupric chloride has maxima at 26.9 per cent and 4 per cent, the latter after the plot has crossed that for cuprous chloride. When cuprous and cupric chlorides are used in combination the curve is approximately a straight line, the effect of the cuprous chloride being inappreciable at high dilutions; incidentally, the curve for the combination is tangential to the peak of the cupric chloride curve. The remarkable increase in activity of cupric chloride at 4 per cent strength suggests the existence of a complex hydrated cupric chloride anion in which the chlorine is much more active than the bound water. This preferential activity of the chlorine may be due to its actual distance from the copper being greater than that of the bound water, as in the known example of crystalline $K_2CuCl_4 \cdot 2H_2O$ (8, 29, 30). The active amount of this complex would appear to be only slowly changed with increasing dilution. (*Cf.* the previous discussion on the non-reactivity of the chlorine in the zinc chloride complexes as possibly due to the water being further from the zinc than the chlorine.)

A further study by Hodgson (33) has shown that while cupric chloride is an effective catalyst for the replacement of the diazo nitrogen in strongly positive diazonium cations (such as are furnished by the nitroanilines) by chlorine, and in both weakly acid and neutral solutions saturated with sodium chloride, it is much less efficient for weak cations such as are given by aniline and the three toluidines, although it is superior to cuprous chloride under these conditions. At the slight acidity of the acid reactions, the catalytic power of the cupric chloride is generally somewhat inferior to that in neutral solutions, a result which confirms an earlier observation of Hodgson and Sibbald (43). In solutions of the less positive diazonium cations the main reaction is phenol formation, and the phenol in neutral solution then tends to couple with undecomposed diazonium compound, a reaction which is very pronounced in benzidine decompositions. It is of interest that neutral solutions of *o*- and *p*-nitrobenzenediazonium chlorides saturated with sodium chloride gave no chloronitrobenzene, and that *m*-nitrobenzenediazonium chloride gave only 6 per cent of *m*-chloronitrobenzene, whereas when hydrochloric acid was present small yields of the chloronitro compounds were obtained in all three cases; these results suggest the viewpoint above (34), that unionized hydrogen chloride is the active reagent and not the chloride ion (44). In similar decompositions with diazotized *o*- and *p*-anisidines there is a surprisingly large replacement of the diazonium group by chlorine, which indicates that the negative ($-I$) inductive effect of the methoxyl group has appreciably increased the positivity of the significant carbon to which the diazonium group is attached.

K. *A limitation of the Sandmeyer reaction as shown by the behavior of diazotized 2,4,6-trichloro- and 2,4,6-tribromoanilines*

In some experiments carried out by Hodgson and Mahadevan (40), it was found that whereas 2,4,6-trichloro- and 2,4,6-tribromoanilines were smoothly diazotized by the nitrosylsulfate–glacial acetic acid procedure of Hodgson and Walker (46) and then behaved normally towards the Sandmeyer reagent (cuprous halide) in concentrated mineral acid solution with formation of tetrahalogenobenzenes, they were reduced in dilute acid (3–4 per cent) to 1,3,5-trichloro- and 1,3,5-tribromobenzene, respectively. In these cases, the cuprous halide has reacted as a reducing agent and not as a catalyst, since cupric chloride or bromide, respectively, was found in the ultimate reaction mixture. This reducing behavior of cuprous salts recalls the somewhat analogous behavior of 2-nitro-1-naphthalenediazonium sulfate towards cuprous hydroxide in glacial acetic–sulfuric acid (39) to form β-nitronaphthalene, when other reaction products might have been anticipated, *viz.*, 2,2′-dinitro-1,1′-azonaphthalene and 2,2′-dinitro-1,1′-dinaphthyl. With potassium iodide, however, both 2,4,6-trichloro- and 2,4,6-tribromobenzenediazonium sulfates gave normal replacement of the diazo group by iodine, whether in concentrated or dilute mineral acid or in neutral solution. Chattaway's observation (6) that potassium perbromide gave 2,4,6-tribromobenzenediazonium perbromide, which on heating with glacial acetic acid gave 1,2,3,5-tetrabromobenzene, has been confirmed.

L. The formation of complex salts between aryldiazonium chlorides and ferric chloride, and the decomposition of p-nitrobenzenediazonium chloride by ferrous salts

It is generally agreed that an unstable intermediate product is formed in the solution during the usual Sandmeyer reaction with cuprous chloride. The work of the author and his collaborators in this field has been based so far on the existence, however ephemeral, of intermediate complex anions; while none have hitherto been actually isolated, ocular demonstration of their existence has been shown (37) for the cases of blue complex anions with cobalt chloride of the type $(RN_2)_2^{++}\}CoCl_4^{--}$. It was then suggested that the efficiency of ferric chloride as a catalyst in concentrated hydrochloric acid was due to the formation of complex salts of the type $R\overset{+}{N}_2\}FeCl_4^-$. Incidentally, Schmidt and Maier (59) have previously obtained double salts of *p*-diethylaminobenzenediazonium chloride with mercurous, ferric, zinc, and cadmium chlorides.

Hodgson and Sibbald have since found (45) that complex salts of aryldiazonium chlorides with ferric chloride of the general formula $R\overset{+}{N}_2\}FeCl_4^-$ were readily prepared, owing to their sparing solubility in concentrated hydrochloric acid. Eight examples of these salts are described from aniline, *m*-nitro- and *p*-nitroanilines, *p*-toluidine, α- and β-naphthylamines, benzidine, and 4,4′-diaminodiphenyl disulfide, a set sufficiently versatile to indicate their ease of formation. These ferri-complexes decomposed when heated with concentrated hydrochloric acid similarly to the original solutions of diazonium chlorides when they were treated with ferric chloride (43), and so furnished strong evidence of intermediate compound formation preceding decomposition in concentrated hydrochloric acid. When the ferrichloride was treated with concentrated hydrobromic acid the chlorine was replaced by bromine, and the ferribromide, $R\overset{+}{N}_2\}FeBr_4^-$, was isolated; it decomposed in hydrobromic acid solution to give excellent yields of RBr.

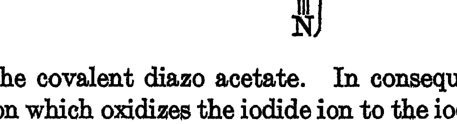
When the ferrichloride from diazotized *p*-nitroaniline was decomposed by boiling hydrobromic acid, a mixture was obtained of *p*-bromonitrobenzene (93 per cent) and *p*-chloronitrobenzene (7 per cent), a decomposition exactly analogous to those already referred to (34) of diazonium salts with cuprous chloride and cuprous bromide in hydrobromic and hydrochloric acid, respectively. The corresponding ferribromide, however, when decomposed by boiling hydrochloric acid gave a mixture of *p*-bromonitrobenzene (50 per cent) and *p*-chloronitrobenzene (50 per cent). As well as substantiating the general interpretation advanced above for the mixed Sandmeyer reactions, the data indicated that in the competitive introduction of chlorine and bromine, where the bromo complex is being decomposed by hydrochloric acid, ferric iron is more favorable to the introduction of bromine than is cuprous copper, since the latter provided a mixture of *p*-bromonitrobenzene (36 per cent) and *p*-chloronitrobenzene (64 per cent).

p-Nitrobenzenediazonium ferrichloride is very rapidly decomposed by ethanol and, while *p*-chloronitrobenzene is the main product, some nitrobenzene is also formed, owing to the usual reduction of diazonium salts by ethanol. In the analogous case of zinc chloride double salts (38), nitrobenzene was the main product (*ca.* 66 per cent) of decomposition with ethanol.

Ferrous chloride and ferrous sulfate also decompose, very efficiently, concentrated hydrochloric acid solutions of *p*-nitrobenzenediazonium chloride, giving *p*-nitrochlorobenzene in about 81 per cent yield; whereas the yield falls only slightly (*ca.* 6 per cent) even at great dilutions, the difference between the two yields when the ferrous salt is present and absent rises appreciably with dilution, thereby proving that the mechanism of decomposition involves the presence of a complex anion containing iron and chlorine.

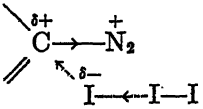
M. The action of iodides on acid solutions of diazonium salts

Experiments have been made (34) to test whether the facile direct replacement of the diazo group by iodine was due to the iodide anion, as was generally supposed, or was bound up with the oxidizing action of the diazo group. For this purpose *p*-nitroaniline was diazotized in hydrochloric acid and treated with sodium bisulfite; *p*-nitrophenylhydrazine hydrochloride then resulted. When, however, the mineral acid was first replaced by acetic acid, sodium bisulfite or sodium thiosulfate was without immediate effect, and subsequent addition of potassium iodide produced no 4-iodonitrobenzene in either case. On addition of the mild oxidizing agent copper sulfate, 4-iodonitrobenzene was immediately precipitated. Further, potassium iodide decomposed a diazo compound much more slowly in dilute acetic acid than in dilute mineral acid, a result which would indicate an equilibrium



$$\mathrm{RN{=}NOCOCH_3} \rightleftharpoons \left.\begin{matrix}\mathrm{R\overset{+}{N}}\\ \equiv\\ \mathrm{N}\end{matrix}\right\}\mathrm{OCOCH_3^-}$$

greatly in favor of the covalent diazo acetate. In consequence, since it is the positive diazonium ion which oxidizes the iodide ion to the iodine radical (neutral atom) initially, the reaction in acetic acid will be slow, whereas in mineral acid, where the diazonium ion is in great excess, reaction will be rapid. By analogy with the action of cuprous salts (as halogen carriers), it is reasonable to assume that the iodine acts also in its familiar rôle as halogen carrier, *via* a mechanism in which the iodine radical first formed by oxidation unites with another to form the molecule I_2, which then by combination with an iodide ion forms I_3^-, and this complex anion reacts with a diazonium ion to form the iodo compound, iodine is liberated, more iodine is regenerated, and so on.



Further, if the reaction between diazonium ion and hydriodic acid were purely the simpler one of straight reduction–oxidation with immediate combination of the aryl and iodine radicals formed, it would be reasonable to expect some diaryl formation (see later), whereas this has not been observed. This interpretation of the direct action of potassium iodide on acid solutions of diazonium salts is strictly analogous to the formation of bromoaryls from aryldiazonium perbromides, $Ar\overset{+}{N_2}\}Br_3^-$, without the intervention of copper as catalyst, but merely by heating with ethanol.

N. Diaryl formation in the Sandmeyer reaction

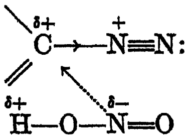
When benzenediazonium chloride was added below 10°C. to a solution of cuprous chloride which had been so diluted with water that the cuprous chloride was just precipitating, the main product was chlorobenzene; but when the sequence of addition was reversed, the chlorobenzene was now accompanied by considerable amounts of diphenyl and 4-hydroxydiphenyl. These experiments indicated that, when the complex anion was in excess, the main attack was at the cationoid carbon to which the diazo group was attached, whereas when the diazonium salt was in excess, a considerable oxidation by the diazonium salt occurred. These experiments also fall into line with the observation of Hodgson and Walker (46), that when *o*- or *p*-nitroaniline was diazotized in sulfuric–acetic acid and treated with a concentrated hydrochloric acid solution of cuprous chloride, i.e., with a minimum of water present, no diaryl formation occurred, whereas under more aqueous conditions considerable quantities of 2,2′- and 4,4′-dinitrodiphenyl were formed. In the former case, the chlorine in the cuprous ion complex was undoubtedly the active agent, while the cuprous chloride functioned solely as a catalyst (halogen carrier); whereas in the latter case, with more water and therefore with more free cuprous chloride present, the copper donates an electron to the diazonium ion, whereby it is oxidized to the cupric ion, nitrogen is evolved, the carbon to which the nitrogen was linked is now in the radical state, and diaryl formation takes place.

The cuprous chloride in the last case was no longer a catalyst but a reducing agent. To test the effect of substituents in the nucleus on diaryl formation, aniline, *p*-bromoaniline, *p*-toluidine, and *o*-nitroaniline were diazotized in hydrochloric acid solution in an atmosphere of carbon dioxide. From the estimation of the cupric chloride formed, the first three indicated 0.2, 0.15, and 0.17 gram-mole, respectively, of diaryl formation per gram-mole of amine taken, while *o*-nitroaniline gave 0.56 gram-mole of 2,2′-dinitrodiphenyl. This indicated that the more positive the diazonium ion, the more readily it was reduced and the copper oxidized. The similar results of *p*-bromoaniline and *p*-toluidine may be ascribed to the mesomeric (resonance) effect (33) of the bromine counteracting its negative ($-I$) inductive effect, since if only the inductive effect had operated, there would have been much greater diaryl formation than that realized.

O. The replacement of the diazonium by the nitro group. A general method analogous to the Sandmeyer reaction based on the decomposition of the aryldiazonium cobaltinitrites (41)

Previous work on the replacement of the diazonium by the nitro group has been carried out by Sandmeyer (55), Veselý and Dvořák (65), Orton (51), Bucherer and van der Recke (4), Hantzsch and Blagden (27), and Contardi (9). In every case, however, the reactions were of limited application (58).

Since diazonium nitrites decompose in nitrous acid solution to form nitro compounds, the mechanism of the reaction would appear to be attack of the



carbon atom to which the diazonium group is attached by anionoid nitrogen (due to resonance) of the nitrite group; otherwise, attack by oxygen would form the hitherto unknown, and probably therefore very unstable, phenyl nitrite, wherein the nitrite group must immediately isomerize into the stable nitro group. This mechanism follows on the analogy of the much greater reactivity of anionoid nitrogen (due to resonance) in nucleophilic nitrous acid (47) than of either of its oxygens for the cationoid carbon atom. Accordingly, complex salts with diazonium cations and metallic nitrite anions were expected to offer a means of preparing nitro compounds by a mechanism analogous to that of the Sandmeyer reaction (34).

For this purpose, the aromatic diazonium cobaltinitrites were prepared by the simple addition of sodium cobaltinitrite to a solution of a diazonium salt previously neutralized by calcium carbonate (41). They were yellow to orange crystalline substances which, when obtained from the nitro- and chloroanilines, were stable when dry; those from the toluidines, anisidines, and α-naphthylamine were less stable. These cobaltinitrites decompose on heating (sometimes with explosive violence), couple with β-naphthol in alkaline solution, and with α- and β-naphthylamines in acetic acid solution, but in hydrochloric acid they diazotize amines and nitrosate phenols by the nitrous acid formed during decomposition of the complex salt by mineral acid. They have the general formula $[RN_2]_3^{+++}\,[Co(NO_2)_6]^{---}$.

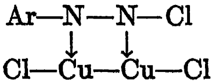
When the solid diazonium cobaltinitrites from aniline, *o*-, *m*-, and *p*-nitroanilines, and *p*-chloroaniline were added to a suspension of cuprous oxide in concentrated aqueous sodium nitrite, a vigorous evolution of nitrogen took place in the cold, and excellent yields of nitro compounds were obtained. However, with the cobaltinitrites from *o*- and *p*-toluidines, *o*- and *p*-anisidines, and

α- and β-naphthylamines, the yields were poor, nor did the further addition of freshly precipitated copper effect any appreciable improvement; when copper sulfate was also present, however, all the cobaltinitrites as yet prepared decomposed to give very favorable yields (mostly above 60 per cent). In the absence of sodium nitrite, cuprous oxide, either alone or in conjunction with copper sulfate, would not decompose the cobaltinitrite in the cold, although on heating there was a reaction without, however, the production of a nitro compound; addition of sodium nitrite to the cold mixture immediately brought about formation of the relevant nitro compound.

The results confirmed the expectations noted above, the anionoid (nucleophilic) nitrite group of the complex cobalt anion attacking the cationoid (electrophilic) carbon atom of the diazonium compound. The simpler mechanism of nitrogen attack, however, would imply that the cobaltinitrite anion contained at least some of its nitrito groups attached to the central cobalt atom by oxygen. This appears a reasonable assumption in view of the facts: (*1*) that the nitrito groups in the metallic cobaltinitrites are indicated by magnetic data to be covalently linked (52); (*2*) by analogy, the mixed inorganic cobalt complex salts containing both NH_3 and NO_2 groups are found in two forms, *viz.*, the red unstable nitritopentammine salts, e.g., $[Co(NH_3)_5(ONO)]X_2$, and the more stable nitropentammines, e.g, $[Co(NH_3)_5(NO_2)]X_2$ (13). Further, both NH_3 and NO_2 are themselves anionoid, but the introduction of NH_3 into the cobalt complex renders it cationoid, whereas the nitrito group renders the complex anionoid. It is reasonable, therefore, to assume that reaction occurs *via* a feebly linked nitrito group, in view of the direct formation of a nitro compound upon the decomposition of a diazonium cobaltinitrite. The function of the copper compounds is to bring about the elimination of the diazonium nitrogen, and the rôle of the sodium nitrite would appear to be the re-formation of stable sodium cobaltinitrite, which would in consequence be the main driving force of the decomposition, since the copper compounds did not themselves effect reaction in the cold.

P. Discussion of Waters' complex cuprous cation theory (71)

Subsequent to the publication by Hodgson, Birtwell, and Walker (34) of their interpretation of the Sandmeyer reaction, Waters (71) suggested a different mechanism for the Sandmeyer reaction and limited that of Hodgson *et al.* as cogent for the thermal decomposition of solid diazonium salts, which, he stated, should be classed as an ionic reaction. The Waters' mechanism, which is almost identical with that for the reduction of a diazonium compound to a diaryl, is based on his statement that Hodgson *et al.* have not been able to suggest why the cuprous salts decomposed in a way which is so very different from the double salts which diazonium halides readily form with the halides of other metals such as zinc, cadmium, mercury, arsenic, antimony, or bismuth which he formulates as $(ArN_2)_2^{++}\}ZnCl^{--}$. He assigns to Hodgson *et al.* the rôle of giving strong support for Hantzsch's view (27) that Sandmeyer reactions are essentially decompositions of complex cuprous diazonium salts formulated by Hantzsch (25) as $(ArN_2,Cu_2)X_3$. It must be pointed out that Hodgson *et al.* did not adopt Hantzsch's formula at all, but emphasized that the catalytic reagent was a cuprous halogeno anion which, in a later paper (35)—submitted for publication nearly three months before the appearance of Dr. Waters' polemic and therefore in entire ignorance of its nature—was definitely formulated by them as CuX^{---}, where X is a halogen atom. It was this cuprous halogeno anion which first formed a salt with the diazo cation, to be followed by the rearrangements already described. Hantzsch's complex had very little in common with the complex anion concept, and it has already been pointed out into what a dilemma Hantzsch would have fallen had he obtained the data which Hodgson *et al.* received from their mixed Sandmeyer reactions. By a curious circumstance, both Hantzsch and Hodgson *et al.* arrived at the same conclusion as to the source of the halogen which replaces the diazo group, though the last-named authors claim validity for their line of argument. Waters himself points out, what is really a criticism of Hantzsch's viewpoint, that while cuprous salts will form addition complexes with covalent azo compounds such as azomethane (11) and for which in consequence it could be suggested that the cuprous diazonium halides are structurally different from all other metal diazonium halides, e.g., that they contain a covalent azo group,



$$\begin{array}{ccccccc} Ar & — & N & — & N & — & Cl \\ & & \downarrow & & \downarrow & & \\ Cl & — & Cu & — & Cu & — & Cl \end{array}$$

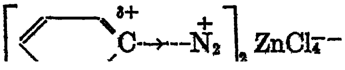
which might decompose to give neutral aryl radicals, nevertheless, the fact that metallic copper can be used instead of a cuprous salt greatly weakens the plausibility of this *ad hoc* suggestion.

The isolation, however, by Hodgson *et al.* of stable complex salts of the type $\overset{+}{RN_2}\}FeCl_4^-$ (45), and $(RN_2)_3^{+++}\}Co(NO_2)_6^{---}$ (41), should leave no possible shadow of doubt as to their conception of what the complex salt would be in the classical Sandmeyer reaction *viz.*, $(RN_2)_3^{+++}\}CuCl_4^{---}$, a conception never put forward either by Hantzsch or by any other author prior to 1941. Further, in reply to Waters' challenge, Hodgson *et al.* (36) reaffirmed emphatically that, far from being unable to suggest why cuprous salts were almost unique as Sandmeyer catalysts, they held the belief (implicit in the very first paragraph of their paper (34)) that there was no justification for such a point of view as that advanced by Waters; in support thereof examples were quoted from both the benzene and the naphthalene series where cupric salts act as efficiently (or nearly so) as cuprous salts, i.e., examples for which the Waters' explanation (71) could have no validity whatever. Moreover, Waters did not mention these reactions, although it is obvious that if cupric salts could behave under identical circumstances like cuprous salts, his whole argument for an oxidation–reduction mechanism involving a complex cation of a cuprous diazonium double salt becomes invalid. In like manner, Sandmeyer's own statement (53) that cupric salts were without the same action as cuprous salts was based on insufficient evidence. Throughout his paper (70) Waters argued for the unique or almost unique character of cuprous salts for replacement of the diazo group by halogens under the conditions of the Sandmeyer reaction, i.e., in aqueous acids, but this unique character does not exist in actual fact (despite Waters' assertion to the contrary, page 268), and the whole argument of a single electron-transfer mechanism from the cuprous copper collapses when the electron comes from the anionoid (nucleophilic) halogen loosely attached to the copper in the complex anion and not from the copper itself. To clinch the argument, it has been shown (33, 43) that below a certain acid strength, cupric chloride is superior as a catalyst to cuprous chloride.

It was pointed out by Hodgson *et al.* in their original paper (34) that the Sandmeyer reaction appeared to be dependent on the degree of positivity of the diazonium ion, which, when of sufficient magnitude, enabled the more ephemeral cupric halogen complexes to react like their more abundant cuprous analogues. This dependence has been subsequently illustrated by the data given above for decomposition by cupric salts, which indicated that when the significant carbon had only a relatively small positive charge (as in diazotized aniline or *p*-toluidine) the yield of chloro compound was only small, but with increase of positivity, as in *p*-chlorobenzenediazonium chloride, and still more so in *p*-nitrobenzenediazonium chloride, owing to the $(-I)$ inductive effect of the chloro and nitro groups, respectively, yields as ample as those with cuprous chloride were obtained. Incidentally, but important for the argument, no diaryl formation was observed in the decompositions with cupric salts, the usual by-product being the corresponding phenol. Waters implicitly concedes this point, since he depicts [71, page 268, Fig. 1] the positive centers of a complex cation of a cuprous diazonium double salt as attracting each other, an attraction which could only result in electron release to the significant carbon of the aryl group when this is at a higher positive potential than the cuprous copper. Hodgson *et al.* suggested that attraction was far more likely to occur between the reactive anionoid center of a complex anion (such as is generally acknowledged to be formed when cuprous salts are dissolved in mineral acids) and the positive significant carbon, than between positive centers in a complex structure such as that depicted by Waters.

Q. Halogen replacements by metallic complexes other than those of copper

A statement by Groggins (20) that copper is not the only metal that can advantageously be employed for the replacement of the diazo group by halogen has already been quoted above. Further, recent work by Hodgson and Foster (38) has established that zinc chloride double salts of diazonium chlorides are decomposed by molten phenols, and, in the cases where a nitro group was not present in the aromatic nucleus of the diazonium chloride, three simultaneous reactions occurred, *viz.*: (*a*) replacement of the diazo group by chlorine, which is the main reaction and affords yields of chloro compounds up to 50 per cent; (*b*) formation of a hydroxydiphenyl; and (*c*) formation of a diphenyl ether. The three different mechanisms involved may be illustrated from the complex involving benzenediazonium chloride,



$$\left[\overset{\delta+}{C}\cdots\!\rightarrow\!\cdots\overset{+}{N}_2\right]_2 ZnCl_4^{--}$$

The cationoid carbon to which the diazo group is attached reacts: (*a*) with anionoid chlorine in the complex, whereby electron release to the carbon occurs with separation of neutral chlorine, evolution of nitrogen, and subsequent linkage of the aryl radical with neutral chlorine to form chlorobenzene in 40 per cent yield, a mechanism analogous to that of the Sandmeyer reaction (34); (*b*) with the anionoid *p*-carbon atom of the phenol to form, by similar release of electrons, etc., the 4-hydroxydiphenyl in 26 per cent yield; (*c*) with the anionoid oxygen of the hydroxyl group of the phenol, whereby the diphenyl ether is produced in 20 per cent yield. The zinc chloride double salts of *o*-, *m*-, and *p*-nitrobenzenediazonium chlorides all coupled with the phenol to form *o*-, *m*-, and *p*-nitrobenzeneazophenols. This enhanced coupling power is due to the well-known stabilizing effect of the nitro on the diazo group. The anionoid (nucleophilic) phenol in the reaction above plays a rôle identical with that of anionoid water.

When the double salts of diazonium chlorides with lead tetrachloride are heated with alcohol, the diazo group is replaced by chlorine (7), the reaction being analogous to the decomposition of the perbromides and indicating the function of the metal to be that of host for the formation of the reactive complex anion.

R. The dual rôle of cuprous compounds

Hodgson *et al.* (36) consider that it is more rational to postulate distinct mechanisms (*1*) for halogen substitution and (*2*) for diaryl formation than a single mechanism such as that proposed by Waters (71) for both. Hodgson and Walker found (46) that when the amount of water was reduced, i.e., the concentration of the complex cuprous ion increased, diaryl formation could be prevented altogether in the decomposition of *o*- and *p*-nitrobenzenediazonium chlorides by cuprous chloride in hydrochloric acid solution. The Waters' mechanism does not explain the complete absence of diaryl formation in such cases, which is due not to the high concentration of chloride ions, since this may be very moderate, but to conditions which favor the maximum production of a complex anion. Likewise, when the water content was increased, the formation of diaryl was also increased but was accompanied by that of nitrophenol, indicating that anionoid water now competed with the cuprous reducing agent for the diazonium salt.

The Waters' mechanism, moreover, differs only in elaboration from that already proposed by Groggins (20), who also introduces hydroxyl anions as competitors with the cationoid complex. Hodgson *et al.* submit that it is more reasonable to assume competition for the diazonium cation between a complex cuprous anion and anionoid water (since there are no hydroxyl ions present in such strong acid solutions, in a realistic sense of the term), than to assume a competing cuprous cation. The fact that by reducing the amount of anionoid copper complex by dilution of the solution with water the yield of halogen compound is decreased, whereas that of phenol is increased, supports this viewpoint.

S. Decompositions by copper sulfate, with and without other oxidizing agents

Reactions between nitroaryldiazonium sulfates, copper sulfate, and sodium chloride or bromide in sulfuric acid solution to give chloro- or bromoaryls in 90–100 per cent yields were discovered by Hodgson (31). Benzenediazonium sulfate itself gives a 38 per cent yield of bromobenzene under similar conditions. To eliminate any chance of cuprous salt formation (in view of Waters' statement as to the unique character of cuprous salts (71)), a number of reactions were carried out with *p*-nitrobenzenediazonium sulfate, copper sulfate, and sodium bromide, in the presence of ammonium persulfate and also of ferric chloride. Although the yields of halogenoaryl were not much affected, nevertheless in the ferric chloride experiment a mixture of about equal parts of 4-chloro- and 4-bromonitrobenzenes resulted. When hydrogen peroxide was used as the supplementary oxidizing agent, the yield of 4-bromonitrobenzene was reduced, and the reaction included the initial formation of a yellow precipitate (probably of a perbromide), which, on heating, decomposed with evolution of bromine and separation of 4-bromonitrobenzene. The last experiment might suggest that when oxidizing agents are present, bromine replacement occurs *via* intermediate formation of a perbromide, but in the decompositions with sodium bromide and copper sulfate alone, the by-product was always the corresponding phenol, which was never found to have been brominated. Further, a mixture of copper sulfate, sodium bromide, and phenol, in the presence of dilute sulfuric acid, gave no precipitate of tribromophenol, even when heated under reflux on the water bath for an hour and subsequently kept for 24 hr. Initial formation of the perbromide does not therefore take place in the copper sulfate and sodium bromide experiments, and the mechanism of decomposition of the diazonium salt is *via* a cupric copper anionoid complex as originally suggested by Hodgson *et al.* (34, 35).

V. CONCLUSIONS

The experimental evidence proves: (*a*) cuprous salts do not possess the unique (or almost unique) character claimed for them first by Sandmeyer (53) and recently by Waters (71). (*b*) In cases where the diazonium cation possesses sufficient positivity, cupric salts can function with efficiencies comparable to those of cuprous salts and, under certain conditions, even superior to them. (*c*) Metals other than copper, such as iron, cobalt, and zinc, can form anionoid complexes with halogens which decompose diazonium cations in like manner to the copper salts. (*d*) The Sandmeyer reaction is an oxidation–reduction process involving an electrophilic diazonium cation and a nucleophilic (anionoid) halogen which is part of a cuprous copper halogeno complex anion, the formula of the intermediate salt, by analogy with the isolable iron salts, $(\mathrm{Ar}\overset{+}{\mathrm{N}})\}\mathrm{FeCl}_4^-$, and with the aryldiazonium cobaltinitrites, $(\mathrm{ArN}_2)_3^{+++}\}\mathrm{Co(NO_2)}_6^{---}$, being $(\mathrm{ArN}_2)_3^{+++}\}$ CuCl_4^{---}. (*e*) There are no fundamental differences in the mechanism of formation of halogen compounds during the decomposition of complex salts of diazonium chlorides with cuprous chloride, cupric chloride, cobalt chloride, ferric chloride, zinc chloride, or any other chloride including hydrogen chloride (similarly for hydrogen fluoride) in aqueous or other media; such differences as do occur, e.g., in the distribution of yield, are due to differences in structure, stability, and concentration of the complex anion, whereby the competing action of anionoid water can become negligible (cuprous salts) or predominant (zinc salts). (*f*) The decomposition of diazonium cations by potassium iodide proceeds *via* the formation of KI_3, and it is the complex anion I_3^- which enters into the mechanism of decomposition, by analogy with the similar decomposition of aryldiazonium perbromides, $\mathrm{Ar}\overset{+}{\mathrm{N}}_2\}\mathrm{Br}_3^-$, and both are in accordance with the mechanism of the Sandmeyer reaction. (*g*) All the reactions discussed in this paper can be interpreted by the simple anionoid mechanism proposed (34, 35). The older theories of Hantzsch, Erdmann, Gattermann, Groggins, etc., are obviously inadequate, and in the cases of Hantzsch, Erdmann, and Groggins, an unnecessary covalent union of copper and nitrogen is assumed.

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THE PYRAZINES

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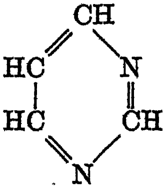
I. INTRODUCTION

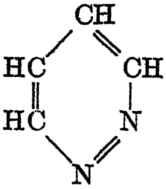
There are three possible ways in which two nitrogen atoms may be combined with four carbon atoms to form the system of six-membered heterocyclic rings known as diazines.

The most familiar class of this system is that in which the two nitrogen atoms are meta to each other, the pyrimidines. These are of importance because of their occurrence in such diverse natural products as the purines (caffeine, uric acid, etc.), vitamins (thiamin, riboflavin), and nucleic acids, as well as in the synthetic barbiturates. T. B. Johnson and coworkers have made extensive investigations of these compounds, and valuable discussions by them may be found in Gilman's *Organic Chemistry* (92) and *Chemical Reviews* (115).



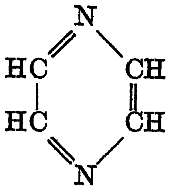
Pyrimidine

The *o*-diazines, the pyridazines, are a more obscure class. For a unified account of their chemistry the reader is referred to Meyer and Jacobson's *Lehrbuch der organischen Chemie.*



Pyridazine

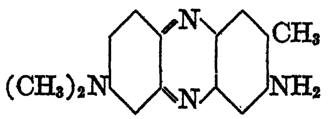
The final class of diazines are the pyrazines, in which the ring nitrogens are situated para to each other. Stoehr and Wolff (*cf.* references) were the first to



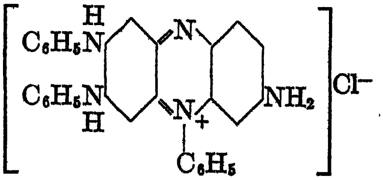
Pyrazine

work intensively in this field, and it is only comparatively recently that further integrated investigations of these compounds were undertaken. This lack of interest was no doubt due to the difficulties which obtained in the synthetic procedures, as well as to the lack of a convenient natural source of supply.

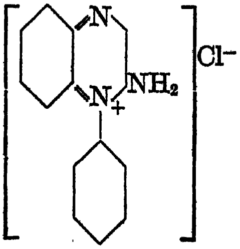
It is worthy of note that Hilditch (103) considered pyrazine to be the most important of the diazines because of its relationship as the parent of those condensed systems which constitute the azine dyes, e.g., the eurhodines, indulenes, and safranines:



Toluylene red
(a eurhodine)

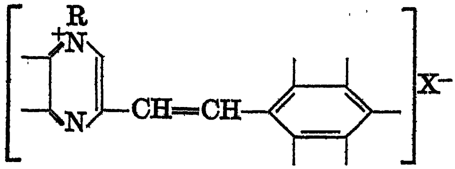


Indulene

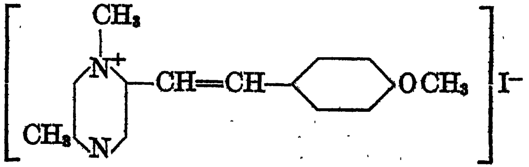


Safranine

More recently (1933), Riesz (182) patented compounds of the type



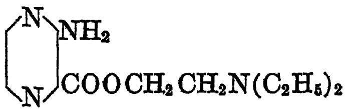
claiming that they possess the property of desensitizing photographic emulsions in aqueous medium. A specific example is 1,5-dimethyl-2-[4′-methoxystyryl]-pyrazinium iodide,

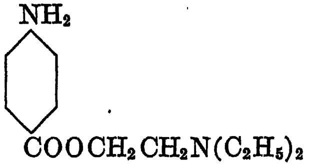


The preparation of pyrazine "phthalein" dyes was recorded by De and Dutta (59) in 1934, when they were successful in condensing *o*-pyrazinedicarboxylic acid with various phenols and amines.

The renewed interest in pyrazine chemistry may be largely traced to the great advances of chemotherapy, in which heterocycles have been particularly prominent. The successful application of sulfapyridine, sulfathiazole, and sulfadiazine led to the development of sulfapyrazine, which has had very favorable clinical reports (132). The indexes of *Chemical Abstracts* for the past four years list over twenty references to clinical investigations and related topics involving sulfapyrazine and its derivatives.

Attempts to improve the action of the procaine type of anesthetic resulted in the synthesis of the β-diethylaminoethyl ester of 2-amino-3-pyrazinoic acid (73):



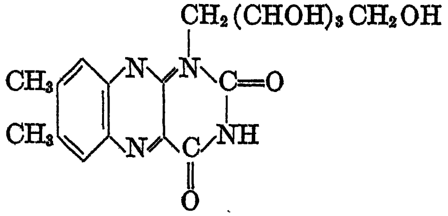


Procaine

The β-diethylaminoethyl ester of 2-amino-5-pyrazinoic acid, which would be more analogous to procaine, was not prepared because of greater difficulties inherent in its synthesis.

Amidopyrazines have been patented by E. Merck (56), with the claim that they are valuable analeptics.

The importance of the pyrazine nucleus in life processes is indicated in its condensed derivative, riboflavin or vitamin B_2.



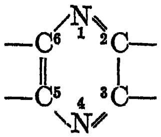
In addition, the simple pyrazinoic and 2,3-pyrazinedicarboxylic acids have both been found to possess antipellagric action (13, 19).

In view of the increasing importance of these substances, the following survey of the chemistry of the *uncondensed* pyrazines is presented with the hope of affording a modern unified picture of this field. We have made no attempts to include the hydro derivatives exhaustively, and these substances are discussed only when they are in some measure concerned with either the synthesis or the reaction of a pyrazine proper.

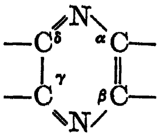
II. NOMENCLATURE

As will be indicated in the following historical section, the development of the nomenclature was beset with difficulties, owing to the confusion of the early workers with regard to the structure of pyrazine. In addition to understandable trivial names given to compounds discovered before their structure was elucidated, there arose such terms for pyrazine as aldine, paradiazine, and piazine. At one time, 2,5-dimethylpyrazine was thought to be the parent of the class, with the result that it was given the generic name "ketine."

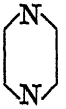
We shall follow the *Ring Index* (169), giving preference to the class name "pyrazine" and numbering the nucleus as follows:



An alternative scheme is that used by Victor v. Richter in his *Lehrbuch der organischen Chemie:*



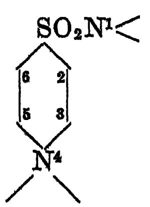
As an abbreviation for the pyrazine nucleus, we shall use the hexagonal skeleton, except where possible ambiguity may occur.



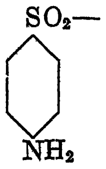
Since the carboxylic acids are among the more widely discussed compounds, it was suggested by Daniels and Iwamoto (57) that for convenience the single monocarboxylic acid be given the name of pyrazinoic acid.

In considering the sulfanilamide derivatives, advantage was taken of the scheme developed by Crossley, Northey, and Hultquist (50) for a systematic nomenclature. For the sake of the unfamiliar reader we quote the more salient features of the system.

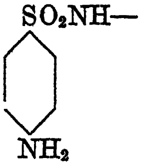
The following numbering sequence is utilized for the substituted sulfanilamides:



The radical names, sulfanilyl

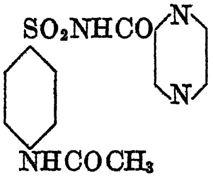


and sulfanilamido

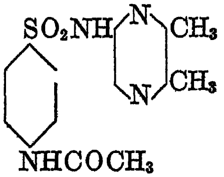


are also found to be very convenient.

We give two examples:



N^4-Acetyl-N^1-pyrazinoylsulfanilamide

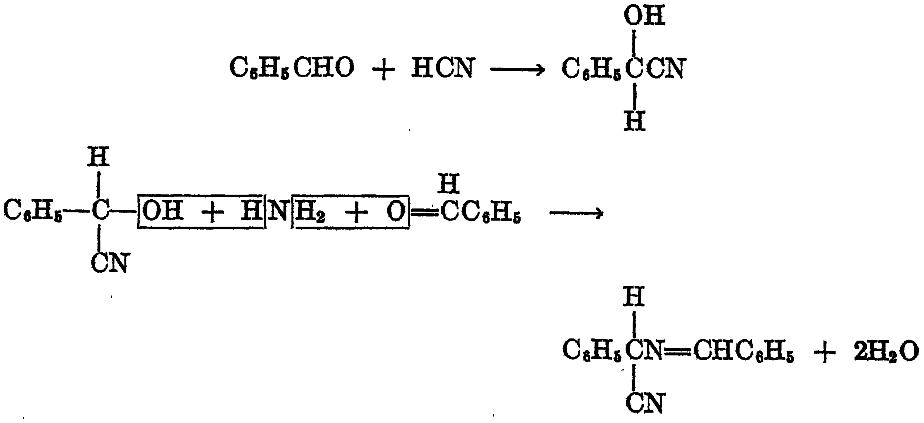


2-[N^4-Acetylsulfanilamido]-5, 6-dimethylpyrazine

III. HISTORICAL

In recording the development of pyrazine chemistry, we have attempted to preserve the original historical plan, presenting the facts in almost the same manner as they were exposed. This treatment may initially appear somewhat awkward, but it will be seen that the early and apparently unrelated observations form an important part of the entire pattern.

The first procedure for the synthesis of a pyrazine derivative was published by Laurent in 1844 (134). Starting with crude benzaldehyde, i.e., benzaldehyde containing some hydrogen cyanide, he treated it with ammonia to obtain what was then known as "benzoylazotid"—actually α-benzalaminophenylacetonitrile:



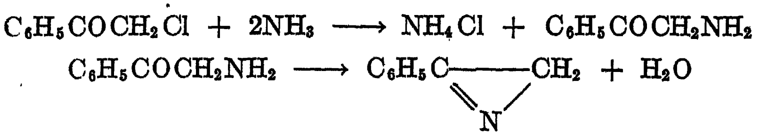
$$C_6H_5CHO + HCN \longrightarrow C_6H_5\overset{OH}{\underset{H}{C}}CN$$

$$C_6H_5-\overset{H}{\underset{CN}{C}}-\boxed{OH + H}N\boxed{H_2 + O}=\overset{H}{C}C_6H_5 \longrightarrow C_6H_5\overset{H}{\underset{CN}{C}}N=CHC_6H_5 + 2H_2O$$

Then, in a fashion typical of that period, he more or less destructively distilled "benzoylazotid," and was able to isolate amongst the products a substance which he called amarone.

Twenty-one years later Erdmann (74) reported an apparently new substance, benzoinimide, which he obtained, along with some others, by the action of ammonia on benzoin.

No further progress was made until 1876, when Stadel and Rügheimer (200) published a paper describing the formation of a new compound, isoindol, by the action of ammonia on ω-chloroacetophenone. They postulated that isoindol was the inner anhydride of an amino ketone and was formed according to the following sequence:



$$C_6H_5COCH_2Cl + 2NH_3 \longrightarrow NH_4Cl + C_6H_5COCH_2NH_2$$

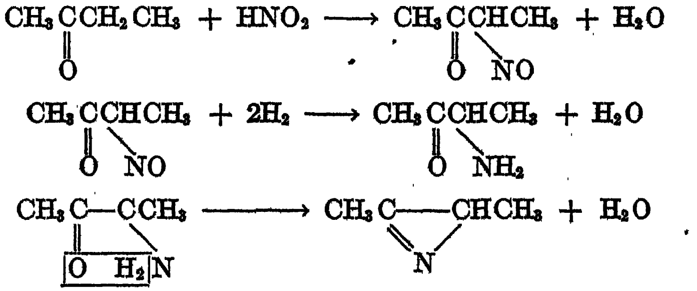
$$C_6H_5COCH_2NH_2 \longrightarrow C_6H_5C\text{—}CH_2\ (\text{N-bridged}) + H_2O$$

Isoindol

This is of interest since it represents the first recorded instance of a structural formula for a pyrazine.

In 1881 Etard (78) made what was seemingly an uncorrelated observation when he published the fact that a new substance, which he named glycolin, could be isolated from the distillate which resulted from heating a mixture of glycerol and an ammonium salt.

Two years prior to the appearance of Etard's article, work was progressing in Victor Meyer's laboratory at Zurich which was to be of great value in the elucidation of the nature of these compounds. At that time Meyer was interested in the study of nitroso derivatives. To H. Gutknecht he assigned the task of reducing the product of the action of nitrous acid on ethyl methyl ketone. We now know that this nitrous acid product was the oxime, but these investigators believed that they were dealing with a true nitroso compound. The results were published in 1879 (99), with the conclusion that an inner anhydride of the amino ketone was obtained from the reduction, in exact analogy to the hypothesis of Stadel and Rügheimer:



$$CH_3\underset{O}{\overset{}{C}}CH_2CH_3 + HNO_2 \longrightarrow CH_3\underset{O}{C}\underset{NO}{C}HCH_3 + H_2O$$

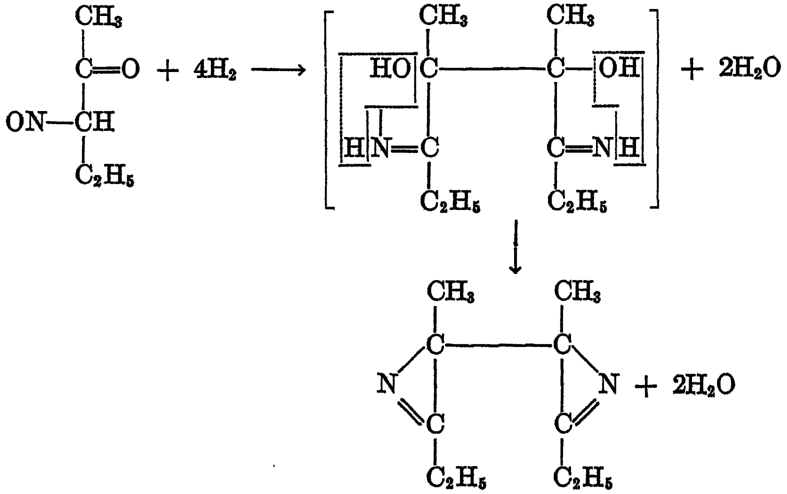
$$CH_3\underset{O}{C}\underset{NO}{C}HCH_3 + 2H_2 \longrightarrow CH_3\underset{O}{C}\underset{NH_2}{C}HCH_3 + H_2O$$

$$CH_3\underset{O}{C}\text{—}\underset{H_2N}{C}CH_3 \longrightarrow CH_3C\text{—}CHCH_3\ (\text{N-bridged}) + H_2O$$

Analyses of the crystalline product were not too satisfactory because of difficulties in obtaining it in an absolutely anhydrous state. After consideration of the data, Gutknecht advanced the possibility that the compound might contain one less hydrogen atom than was indicated by the formula.

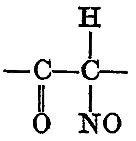
This point was finally clarified by F. P. Treadwell (of analytical fame) in the same laboratory in 1881 (221). He had carried out a reduction of (iso) "nitrosoethylacetone" and had isolated a crystalline hydrate. After placing these crystals in a desiccator over calcium chloride he noted that an anhydrous oil had formed. Analysis of this oil proved that it possessed one hydrogen atom less than would be anticipated from the formula of the simple "inner anhydride" of an amino ketone. Vapor-density determinations further indicated that the molecular weight was about twice that which would be expected for such a compound.

Using the reduction of acetone to pinacol as a model, Treadwell assumed that the reaction for (iso) "nitrosoethylacetone" proceeded as follows:

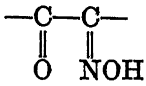


The name "ketine" was applied to this new series of nitrogenous compounds to indicate their derivation from ketones. The simplest member, according to the Treadwell theory, would be that obtained from acetone; and this was specifically termed ketine. The other members were named as derivatives of ketine, so that Gutknecht's compound was called dimethylketine, and Treadwell's diethylketine.

The next year V. Meyer submitted a paper (151) in which he contributed two important suggestions. He first pointed out that although the products of the action of nitrous acid on the ketones were presumably nitroso compounds, they failed to respond to the Liebermann nitroso test. This led to the hypothesis that they were not true nitroso compounds but rather the isomeric oximes:



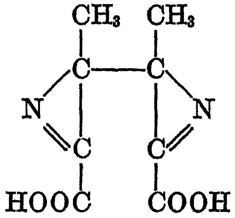
Nitroso ketone



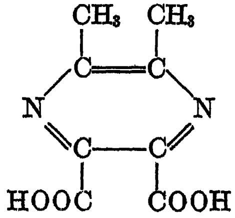
Oximino ketone

He then considered the reduction of these compounds. After making careful comparisons with the corresponding processes for nitro–amine and ketone–pinacol reductions, he concluded that the former was more analogous. This led to the abandonment of the Treadwell theory.

The first conception of the ketines as ring compounds was put forth, almost immediately, by Wleugel in an article concerned with the reduction of (iso) "nitrosoacetoacetic" ester (234). Thus, instead of writing the structure for the "ketine" which was obtained as

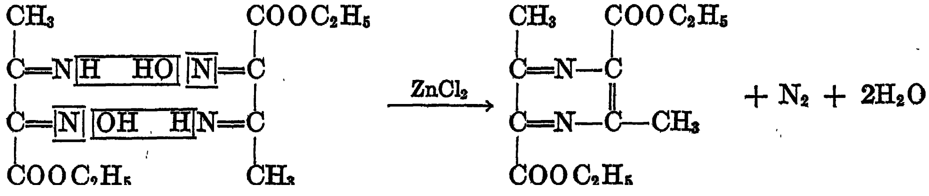


he closed the six-membered ring by utilizing the C—C bonds involved in the two inner anhydride rings:



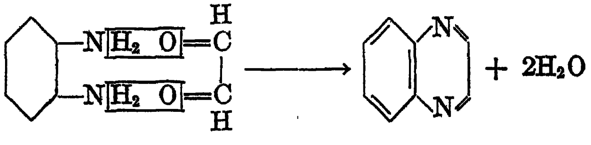
This resulted in the formulation of a heterocycle which, as Wleugel stated, could be conceived to be a pyridine in which the CH group para to the nitrogen atom was replaced by another nitrogen atom. This is our modern concept of the pyrazine nucleus; but there remained a flaw in its derivation, since Wleugel had still utilized the ketine(pinacol) mechanism.

It remained for L. Oeconomides in 1886 (167) to demonstrate experimentally that this mechanism was untenable. He attempted to dehydrate Wleugel's diacid to the acid anhydride, a reaction which should have been clearly possible if the two carboxyls were ortho to one another. This was unsuccessful, and the natural conclusion was that these functional groups had been assigned to incorrect positions on the ring. Verification came, together with a proof that the carboxyls were actually para, from the following experiment. Iminoisonitrosobutyric ester was heated with fused zinc chloride. An examination of the only plausible mechanism which could yield a "ketine" indicated that the carboxyl groups in such a compound would unambiguously be situated at the para positions:

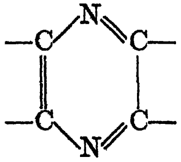


A small amount of free acid was isolated and compared with Wleugel's; the two were found to be identical.

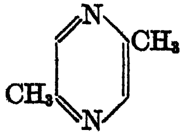
Oeconomides further called attention to the fact that Hinsberg (105) had synthesized quinoxaline, a condensed pyrazine, from *o*-phenylenediamine and glyoxal:



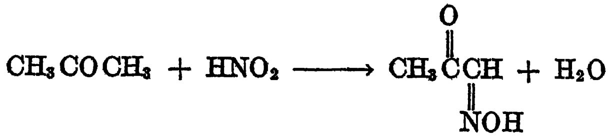
Thus, the ketine nucleus was firmly established as



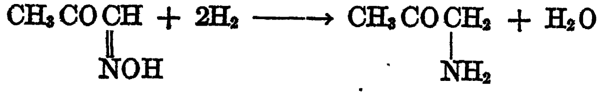
and what was previously thought to be the simplest member, ketine, in fact was the dimethyl derivative.



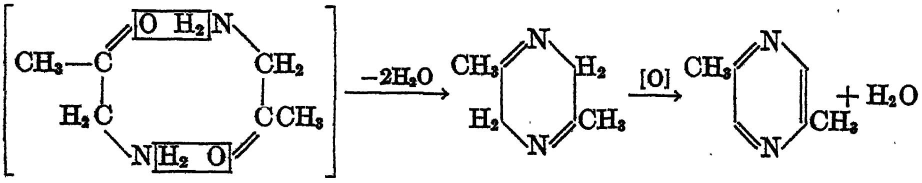
The name "pyrazine" was independently suggested for the nucleus in the following year by Mason (145) and Wolff (236), in order to point up the correlation with pyridine. It is interesting to note that, in the same paper, Wolff acknowledged that the mechanism, first inferred by Meyer, for the preparation of a pyrazine by the reduction of an isonitroso ketone involved an intermediate amino ketone which immediately condensed with itself to yield a dihydropyrazine that was oxidized to the desired pyrazine. Thus, starting with acetone:



$$CH_3COCH_3 + HNO_2 \longrightarrow CH_3\overset{\overset{O}{\|}}{C}\underset{\underset{NOH}{\|}}{CH} + H_2O$$



$$CH_3CO\underset{\underset{NOH}{\|}}{CH} + 2H_2 \longrightarrow CH_3CO\underset{\underset{NH_2}{|}}{CH_2} + H_2O$$

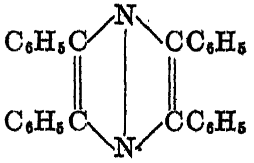


V. Meyer (152) objected to the term "pyrazine" on the grounds that Knorr (126) had already used it for pyrazole tetrahydride, and, in its stead, proposed the generic name "aldine," since the simplest member would result from the self-condensation of the hypothetical aminoacetaldehyde.

Widman (228) finally resolved the issues with a systematic nomenclature. He classified as azines those compounds which contained a six-membered ring consisting of nitrogen and carbon atoms. Hence, substances containing two nitrogen atoms in the ring were called *di*azines. These were further classified, according as to whether the nitrogens were ortho, meta, or para, as *o*-diazines, *m*-diazines, or *p*-diazines. Mason condensed these names, respectively, into oiazines, miazines, and piazines; but it seems that he and his associates were the only ones to use this terminology consistently.

In the light of the newly elucidated structures of the pyrazines, the results of the early workers were finally clarified. Thus, Wolff in 1887 demonstrated that what Stadel and Rügheimer had called isoindol was 2,5-diphenylpyrazine. Stoehr (202) improved Etard's procedure for glycolin and showed that this was identical with 2,5-dimethylpyrazine.

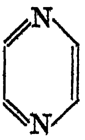
Japp and Wilson (114) undertook the task of determining the true nature of Erdman's benzoinimide, which they renamed ditolanazotide, by repeating the original experiments. In the following year Japp and Burton (110) concluded that the compound was an azine, and assigned to it the structure,

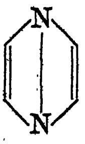


together with the name "tetraphenylazine."

Snape and Brooke thereupon repeated Laurent's original work and in 1897 published a paper (195) in which they revealed that amarone, benzoinimide, ditolanazotide, and tetraphenylazine were all one and the same substance: tetraphenylpyrazine.

There yet remained doubt as to the exact location of the double bonds in the pyrazine molecule. The Kekulé type with its conjugated double bond system (I) and the Dewar type (II) with the long para bond, each had its adherents (147, 212, 246).

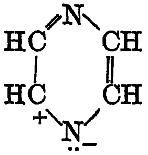

I


II

Bruhl (32) finally established the validity of I after a study of the molecular refractions of a number of pyrazine derivatives.

Recently, Pauling and his collaborators (170) have made electron-diffraction studies on some cyclic systems, including pyrazine. Their results showed that pyrazine and benzene have almost identical structures. Further, although the C—C distances for the two molecules were almost identical, 1.39 A., the value for the C—N distance in pyrazine, 1.36 A., was greater than expected for the Kekulé resonance, which would give this bond 50 per cent double-bond character.

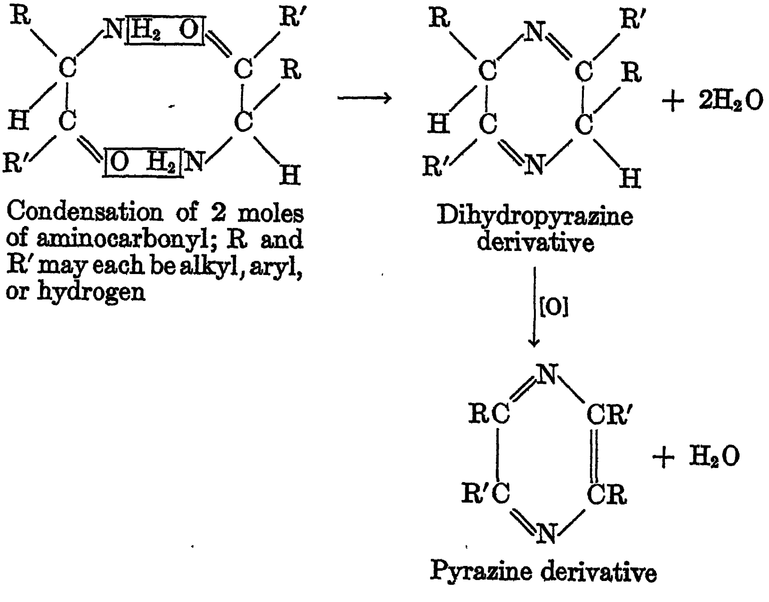
This is interpreted as being due to the large electronegativity of the nitrogen atom, which results in the introduction of additional resonating ionic structures, as:



IV. Methods of Preparation

A. AUTOCONDENSATION OF α-PRIMARY AMINOCARBONYLS

The most general procedure for the synthesis of symmetrical pyrazines depends on the fact that α-primary aminocarbonyl compounds can rarely be isolated because of their strong tendency to condense and form dihydropyrazines. Thus, although the ammonium salts, such as the hydrochlorides, of these compounds are perfectly stable and may be obtained as crystalline solids, most attempts to isolate the free bases by treating with excess alkali yield only the dihydropyrazines. These are easily oxidized by such mild reagents as mercuric or cupric ions, or even air, to the desired pyrazines.



Condensation of 2 moles of aminocarbonyl; R and R′ may each be alkyl, aryl, or hydrogen

Dihydropyrazine derivative

Pyrazine derivative

It is clear that the essence of this method lies in the preparation of the pertinent α-primary aminocarbonyl derivatives. A detailed discussion of these compounds is not inappropriate here, since they are of great importance in pyrazine chemistry. It may be pointed out that in a recent lengthy review paper dealing with amino ketones (49), the primary α-amino type received only cursory mention.

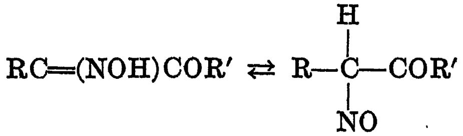
1. *Reduction of oximino ketones*

The first step of this important classical synthesis, discovered by Gutknecht (99), generally involves the formation of the oximino ketone from the ketone:

$$RCH_2COR' + HNO_2 \rightarrow RC{=}NOHCOR' + H_2O$$

R = alkyl, aryl, or hydrogen; R′ = alkyl or aryl.

These oximino compounds were improperly referred to by the early workers as isonitroso ketones on the assumption that an equilibrium existed between these and the true nitroso compounds:



$$RC{=}(NOH)COR' \rightleftarrows R{-}\overset{\overset{\displaystyle H}{|}}{\underset{\underset{\displaystyle NO}{|}}{C}}{-}COR'$$

However, in most cases the species on the right is non-existent (192). Wherever we have used the prefix "isonitroso," it is to avoid possible confusion when reference is made to original papers.

The most direct approach utilizes the action of free nitrous acid on the ketone. This is most often accomplished by dissolving the ketone in acetic acid and then slowly adding, with agitation, a concentrated aqueous solution of sodium nitrite to the well-cooled reaction mixture. The oximino ketone may then be extracted with ether to effect its separation; however, it has been found in our laboratories (15) that a great improvement in manipulation and yield resulted when no attempt was made to isolate the compound for the next step in the synthesis.

Unless the ketones are β-diketones or β-ketonic esters, the reaction may be unsuccessful. Recourse may then be had to the method of Claisen and Manasse (42), which has been found to work in every case. Instead of sodium nitrite, a nitrous acid ester, usually amyl nitrite or occasionally methyl or ethyl nitrite (193), is used in conjunction with either an alkaline catalyst such as sodium ethylate, or an acid catalyst such as hydrogen chloride.

Certain of the oximino compounds may be prepared in better yield by the procedure of Charrier (38) and Freon (86). A β-ketonic ester is dissolved in a cold caustic solution and permitted to hydrolyze to the alkali salt. The mixture is then treated with sodium nitrite and acid to form the oximino ketonic acid, which is unstable and loses carbon dioxide. Oximinoacetone prepared in this way was obtained in 95 per cent yield, as compared to the 40 per cent yield from the Claisen method with acetone (135).

The reduction of the oximino ketone ordinarily is effected in acid medium with either stannous chloride or zinc:

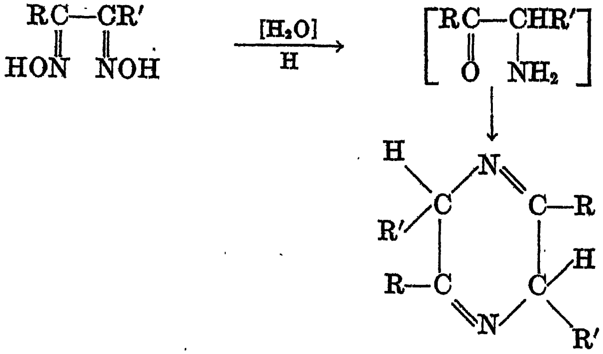
$$RC(=NOH)COR' + 2H_2 + H^+ \rightarrow RCH\overset{+}{N}H_3COR' + H_2O$$

Other reduction techniques have been recorded. Cerchez and Colesiu (39) used an aluminum amalgam to reduce oximinoacetoacetic ester. Satisfactory catalytic reductions were carried out by Diels and Poetsch (63) with a palladium catalyst and by Adkins and associates (3, 229, 230) using Raney nickel. The latter further observed that the yield of pyrazines was a function of the hydrogen pressure, higher pressures favoring the formation of the aminohydroxy derivatives. This is interpreted to mean that at higher pressures there is not sufficient time for the self-condensation to occur, since reduction to the hydroxyamines is so much more rapid. It has been reported that the conditions of the catalytic reduction may be modified so as to favor the formation of piperazines (94).

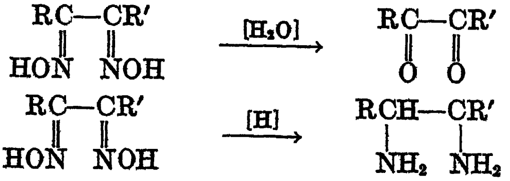
With catalytic hydrogenation the reduction to the aminopyrazine takes place directly, whereas in the acid reduction process the product is the salt of the amino ketone. In the latter case treatment with alkali is necessary to release the free amino ketone, which then undergoes condensation.

The oxidation of the dihydropyrazine is simple and proceeds with great ease. Customarily mercuric oxide or chloride or copper sulfate has been used, but in some cases even atmospheric oxygen has been found to be sufficient.

A number of examples have been recorded where the reduction of diketone dioximes has yielded pyrazine derivatives (table 1). There are two possible ways of picturing the mechanism. Either we may consider that simultaneous hydrolysis and reduction occur in the same molecule:



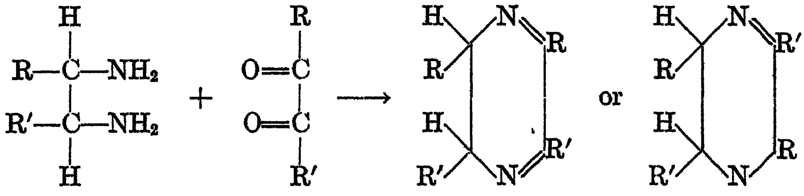
or one molecule undergoes complete hydrolysis and another complete reduction:



$$\begin{array}{c}\mathrm{RC{-}CR'}\\ \|\quad\|\\ \mathrm{HON\ \ NOH}\end{array} \xrightarrow{[H_2O]} \begin{array}{c}\mathrm{RC{-}CR'}\\ \|\quad\|\\ \mathrm{O\quad O}\end{array}$$

$$\begin{array}{c}\mathrm{RC{-}CR'}\\ \|\quad\|\\ \mathrm{HON\ \ NOH}\end{array} \xrightarrow{[H]} \begin{array}{c}\mathrm{RCH{-}CR'}\\ |\qquad|\\ \mathrm{NH_2\ \ NH_2}\end{array}$$

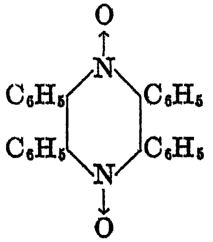
followed by condensation of the two different molecules:



(If R = R′ only one possibility arises.)

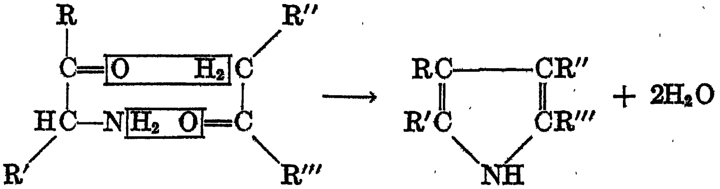
It should be possible to distinguish between the two mechanisms from a study of the dihydropyrazines, since a different one would be obtained in each case; but so far as is known, this task has as yet not been undertaken.

An interesting variation was developed by Durio and Bissi (66), based on the treatment of α-benzil dioxime with alkaline ferrocyanide. This experiment was first performed by von Auwers and Meyer (9), who were able to isolate diphenyl peroxide and a small quantity of an unidentified compound. Durio and Bissi demonstrated that this substance was dioxotetraphenylpyrazine:



and were able to modify the conditions so as to improve its yield. They also found that the other isomers of benzil dioxime were capable of forming only diphenyl peroxide; and, further, that the reaction was not general for the α-aryl dioximes, since α-*p*-tolyl dioxime and α-anisil dioxime yielded no pyrazines.

Before closing our discussion of the reduction of the oximino ketones we should indicate its value in the Knorr pyrrole synthesis (125). The latter is carried out in essentially the same manner, with the exception that after the reduction of the oximino ketone, an equimolar quantity of a ketone (or, better, a β-diketone) is added before the mixture is made alkaline:



It is easily perceived how pyrazines may be obtained as by-products from this reaction, since there is always the possibility of autocondensation of some

TABLE 1

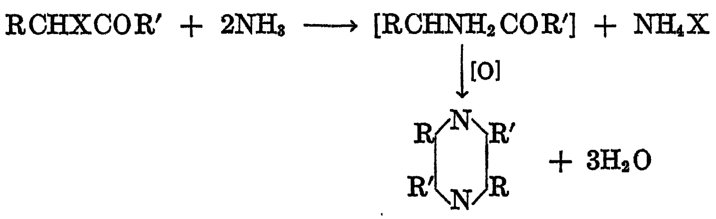
Pyrazines prepared by reduction of oximino ketones

OXIMINO KETONE	PYRAZINE DERIVATIVE	REFERENCES
Oximinoacetone	2,5-Dimethylpyrazine	(4, 87, 106, 114 223)
Oximinoacetoacetic ester	2,5-Dicarbethoxy-3,6-dimethylpyrazine	(3, 39, 88, 97, 230 234)
Oximino ethyl methyl ketone	Tetramethylpyrazine	(26, 28, 78, 99, 165
Oximino methyl propyl ketone	2,5-Diethyl-3,6-dimethylpyrazine	(88, 221)
Oximino methyl butyl ketone	2,5-Dimethyl-3,6-dipropylpyrazine	(131, 222)
Oximino isoamyl methyl ketone	2,5-Dimethyl-3,6-diisobutylpyrazine	(133)
Oximino *n*-hexyl methyl ketone	2,5-Dimethyl-3,6-diamylpyrazine	(14)
Oximinoacetophenone	2,5-Diphenylpyrazine	(29, 172, 230)
Benzyloximinoacetone	2,5-Dibenzyl-3,6-dimethylpyrazine	(196)
Benzil monoxime	Tetraphenylpyrazine	(30, 230)
Dimethylglyoxime	Tetramethylpyrazine	(230)
Acetylbenzoyl dioxime	2,5-Dimethyl-3,6-diphenylpyrazine	(127)
β-Hydroxy-γ-oximino-*n*-valeric acid	Tetramethylpyrazine	(242)
Benzaldiacetyl monoxime	2,5-Dimethyl-3,6-di[β-phenylethyl]pyrazine	(63)
Phenyloximinoacetone	2,5-Dimethyl-3,6-diphenylpyrazine	(218)
Anisyloximinoacetone	2,5-Dimethyl-3,6-dianisylpyrazine	(218)
α-*p*-Tolyloximinoacetone	2,5-Dimethyl-3,6-di-*p*-tolylpyrazine	(218)
Oximinoacetylacetone	2,5-Dimethyl-3,6-diacetylpyrazine	(84, 164)
Benzil dioxime	Tetraphenylpyrazine Dioxotetraphenylpyrazine	(80, 176, 229)
Oximino-3,4-dimethoxyphenyl-3,4-methylenedioxy-β-phenyl ethyl ketone	2,5-Di[3,4-dimethoxyphenyl]-3,6-di[3,4-methylenedioxyphenyl]-	(36)

of the amino ketone. Thus, in the preparation of 3,5-diacetyl-2,4-dimethylpyrrole, some 3,6-diacetyl-2,5-dimethylpyrazine was obtained (84). Indeed, in some circumstances the formation of pyrazines is favored, as was shown by Ochiai, Tsuda, and Ikuna, when they found much tetramethylpyrazine after a reduction of oximino ethyl methyl ketone in the presence of ethyl 2-picoloylacetate (165).

2. Replacement of the halogen in a halocarbonyl with an amino group

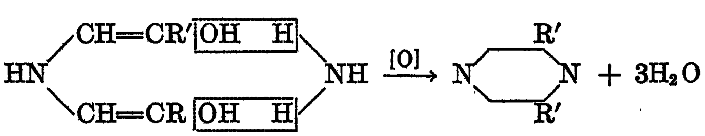
Stadel and Rügheimer (200) were the first to use this method, which has since proved of wide utility.



$$RCHXCOR' + 2NH_3 \longrightarrow [RCHNH_2COR'] + NH_4X$$

R' and R = alkyl, aryl, or hydrogen.

Ammonia, either aqueous or anhydrous, has been employed most often to effect the replacement of the halogen. This, of course, necessitates the use of sealed tubes or autoclaves, since the reaction normally proceeds at elevated temperatures. It might be anticipated that the use of ammonia would give rise to secondary and tertiary amino derivatives: $HN(RCHCOR')_2$ and $N(RCHCOR')_3$. However, the cyclization of the primary aminocarbonyl in most cases is much more rapid. The occasional occurrence of a 2,6-derivative (224) may be explained by the formation of such a secondary amino derivative, which reacts further:



The use of ammonia may lead to other complications, such as resin formation. This was found true in the attempted conversions of *p*-methoxy-ω-chloroacetophenone (224) and α-chloromethyl ethyl ketone (128) to the corresponding pyrazines. In such cases the amination may successfully be accomplished by the Gabriel phthalimide method.

It is found true here, as in other homologous series, that the application of a general method to the first member leads to anomalous results. Many attempts have been made to obtain pyrazine by the action of ammonia on a haloacetaldehyde, but the yields have always been poor (16, 40, 87, 97, 201). Chichibabin and Schukina (40) made a careful study of the conditions and concluded that the

TABLE 2

Pyrazines prepared by the replacement of the halogen in a halocarbonyl with an amino grou

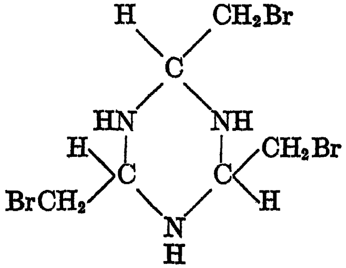
HALOCARBONYL	PYRAZINE DERIVATIVE	REFERENCES
ω-Chloro(or bromo)acetophenone	2,5-Diphenylpyrazine; also some 2,6-diphenylpyrazine	(27, 172, 200, 224 239)
Methyl α-chloroethyl ketone	Tetramethylpyrazine	(61)
Methyl α-chloropropyl ketone	2,5-Dimethyl-3,6-diethylpyrazine	(128)
α-Chlorodiethyl ketone	2,5-Dimethyl-3,6-diethylpyrazine	(61)
β-Bromolevulinic acid	Tetramethylpyrazine	(237)
Bromophenylacetone	2,5-Dimethyl-3,6-diphenylpyrazine	(127)
Phenyl α-bromoethyl ketone	2,5-Dimethyl-3,6-diphenylpyrazine	(43, 188)
Phenyl α-bromopropyl ketone	2,5-Diethyl-3,6-diphenylpyrazine	(43)
ω-Chloro-*p*-methoxyacetophenone	2,5-Di-*p*-methoxyphenylpyrazine (also some 2,6-di-*p*-methoxyphenylpyrazine)	(224)
ω-Chloro-*p*-hydroxyacetophenone	2,5-Di-*p*-hydroxyphenylpyrazine (also some 2,6-di-*p*-hydroxyphenylpyrazine)	(224)
ω-Dibromoacetophenone	3,6-Diphenyl-2-hydroxypyrazine (isoindileucin)	(71)
α-Chloromethyl ethyl ketone	2,5-Diethylpyrazine	(128)
Indacyl bromide	2,5-Diindolylpyrazine	(184)
Bromoacetylskatole	Bis(3-methyl-2-indolyl)pyrazine	(184)
α-Methylindacyl chloride	Bis(2-methyl-3-indolyl)pyrazine	(183)
ω-Chloro-*m*,*p*-dimethoxyacetophenone	2,5-Di[m,*p*-dimethoxyphenyl]-pyrazine	(224)
ω-Chloro-*o*,*p*-dihydroxyacetophenone	2,5-Di[*o*,*p*-dihydroxyphenyl]-pyrazine	(224)

reaction is a function of the temperature, the solvent, and the speed with which the ammonia is passed into the reaction mixture. In addition, they indicated that it may proceed in two directions. The ammonia may add to the carbonyl group to yield the aldehyde ammonia, which loses water to form an imine that polymerizes into a tri[bromomethyl]hexahydrotriazine:

$$BrCH_2CHO + NH_3 \longrightarrow BrCH_2CHOHNH_2$$

$$BrCH_2CHOHNH_2 \longrightarrow BrCH_2CH{=}NH + H_2O$$

$$3BrCH_2CH{=}NH \longrightarrow$$



On the other hand, the bromine of the bromoaldehyde may be replaced by ammonia to form a mixture of various amino derivatives. The best yield of pyrazine, after oxidation of the reaction mixture with mercuric salts, was about 16 per cent.

TABLE 3

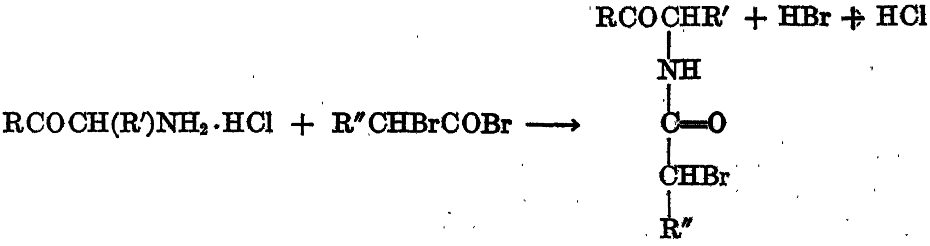
Pyrazines prepared according to the scheme of Tota and Elderfield (220)

R	R′	R″	PYRAZINE
CH_3	CH_3	H	2,3-Dimethyl-5-hydroxypyrazine
C_6H_5	CH_3	H	2-Phenyl-3-methyl-5-hydroxypyrazine
C_6H_5	CH_3	CH_3CH_2	2-Phenyl-3-methyl-6-ethyl-5-hydroxypyrazine
CH_3	$CH_2CH_2OC_2H_5$	H	2-Methyl-3(β-ethoxyethyl)-5-hydroxypyrazine

A more effective method might be to hydrolyze aminoacetal, which may be prepared either from chloroacetal and ammonia (235, 240) or by the reduction of nitroacetal (143). Woodward and Doering (249) improved the former procedure and were able to obtain 72.5 per cent of aminoacetal in addition to 10 per cent of diacetalylamine.

Tota and Elderfield (220) have recently reported a general synthesis for 2,3-disubstituted or 2,3,6-trisubstituted 5-hydroxypyrazines (see table 3).

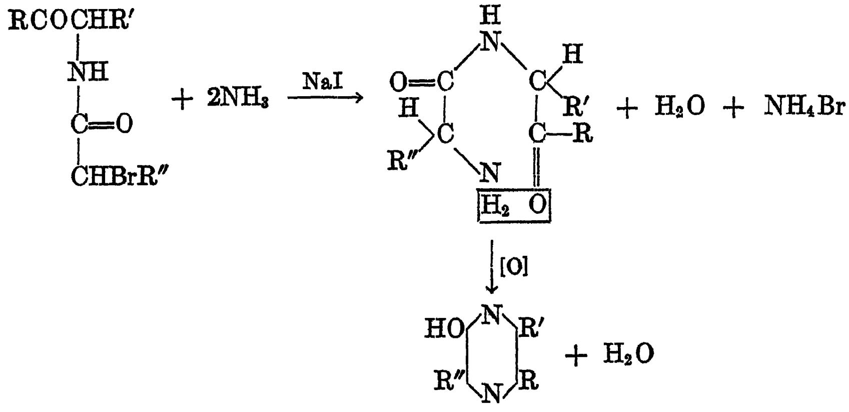
The hydrochloride of a primary α-amino ketone is first treated with a bromoacyl bromide. The more reactive acyl halogen attacks the amino group to form the amide derivative:



$$RCOCH(R')NH_2 \cdot HCl + R''CHBrCOBr \longrightarrow RCOCHR'(NHCOCHBrR'') + HBr + HCl$$

R,R′ = alkyl or aryl; R″ = H or alkyl.

The second bromine atom is then replaced with an amino group by ammoni in the presence of sodium iodide. The resultant molecule then cyclizes by th elimination of water to form a dihydropyrazine, which is oxidized by atmospheri oxygen to the pyrazine:



3. *Reduction of amino acids*

Neuberg (161) reduced the hydrochlorides of the ethyl esters of glycine and alanine to the salts of the corresponding aminocarbonyls with sodium amalgam in acid medium. Subsequent treatment with excess alkali and mercuric chloride yielded fair amounts of pyrazine and 2,5-dimethylpyrazine.

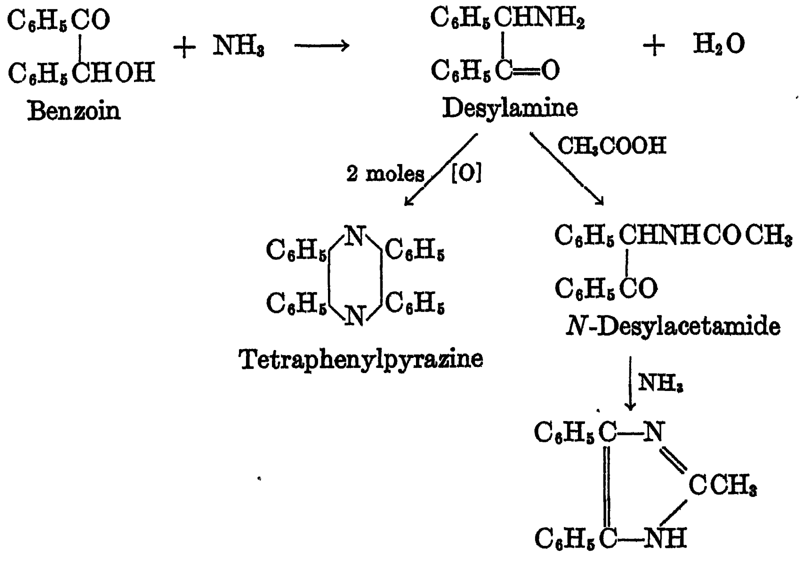
4. *Oxidation of amino alcohols*

Aston and his collaborators used this approach for the synthesis of pyrazine (6). Ethanolamine was passed over a copper catalyst at 300°C. to effect dehydrogenation. The aminoacetaldehyde produced cyclized to dihydropyrazine, which underwent oxidation to pyrazine. The best yield of several experiments was 5.6 per cent, most of the aminoacetaldehyde being consumed by side reactions to form resins. The latter further interfered by inactivating the catalyst.

5. *Replacement of the hydroxyl in a hydroxycarbonyl with an amino group*

This method was first applied by Erdmann (74) for the synthesis of tetraphenylpyrazine by the action of ammonia on benzoin in a sealed tube at elevated temperatures. Japp and Wilson (114) reported a 55 per cent yield with fused ammonium acetate.

Davidson, Weiss, and Jelling (58) simply refluxed benzoin and ammonium acetate in glacial acetic acid. In addition to 20–30 per cent of tetraphenylpyrazine, they obtained equivalent quantities of 2-methyl-4,5-diphenylglyoxaline. To account for these facts, they proposed the intermediate formation of desylamine:

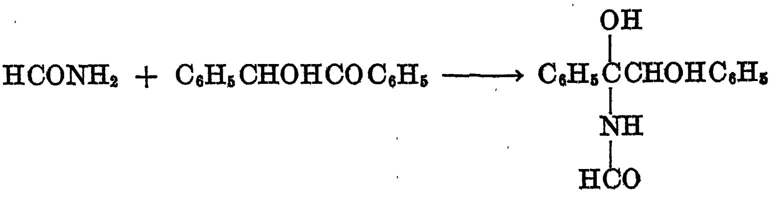


2-Methyl-4,5-diphenylglyoxaline

In support of this mechanism, they demonstrated that replacing benzoin by an equivalent quantity of desylamine hydrochloride gave identical results. Furthermore, although the substitution of propionic acid for acetic acid did not influence the yield of tetraphenylpyrazine, it did lead to the formation of a different imidazole: 2-ethyl-4,5-diphenylglyoxaline. However, when formic acid was the solvent, no pyrazines could be detected.

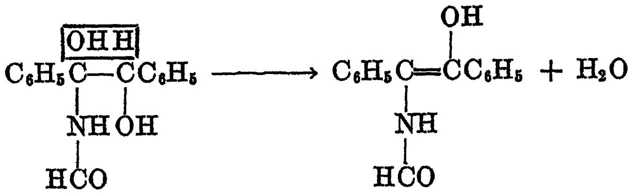
Leuckart (137) had reported that by heating a mixture of ammonium formate and benzoin at 230°C. (quite different conditions from Davidson's) he obtained quantitative conversions to tetraphenylpyrazine.

Ingersoll *et al.* (108) had shown that formamide was the active agent in the formation of primary amines from ketones heated with ammonium formate. However, when Novelli (162) attempted to apply this result to the Leuckart reaction, by heating formamide with benzoin or its derivatives, he could only obtain about 10 per cent of the tetraarylpyrazines. In contrast, this reaction simultaneously produced about 80 per cent of glyoxalines. The following mechanism was suggested. First, there occurs the formation of an unstable addition compound:



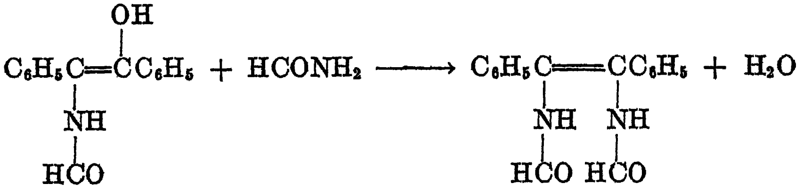
$$HCONH_2 + C_6H_5CHOHCOC_6H_5 \longrightarrow C_6H_5\overset{\displaystyle OH}{\underset{\displaystyle NH\,HCO}{C}}CHOHC_6H_5$$

followed by loss of water:



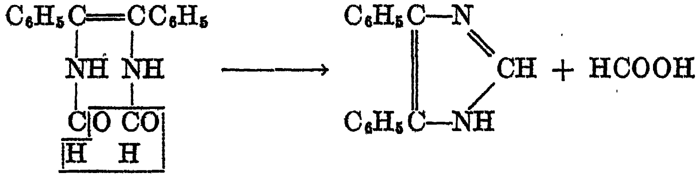
$$C_6H_5\underset{\substack{|\\NH\\|\\HCO}}{\overset{\boxed{OH}}{C}}-\underset{\substack{|\\OH}}{\overset{\boxed{H}}{C}}C_6H_5 \longrightarrow C_6H_5\underset{\substack{|\\NH\\|\\HCO}}{C}=\overset{\substack{OH\\|}}{C}C_6H_5 + H_2O$$

The product reacts with another mole of formamide:

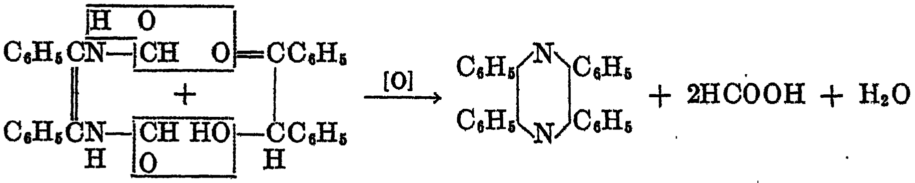


$$C_6H_5\underset{\substack{|\\NH\\|\\HCO}}{C}=\overset{\substack{OH\\|}}{C}C_6H_5 + HCONH_2 \longrightarrow C_6H_5\underset{\substack{|\\NH\\|\\HCO}}{C}=\underset{\substack{|\\NH\\|\\HCO}}{C}C_6H_5 + H_2O$$

4,5-Diphenylglyoxaline (80 per cent) then results by the elimination of formic acid:



Pyrazines are formed (10 per cent) by condensation with a mole of unchanged benzoin.



It is now interesting to speculate why Davidson and his students found no pyrazines when operating in formic acid. There is little doubt that under their experimental conditions very little formamide would be formed, thus invalidating Novelli's hypothesis. On the other hand, the fact that 4,5-diphenylglyoxaline was obtained would indicate the presence of the intermediate desylamine, in accordance with their mechanism. However, it would also be anticipated that some autocondensation would take place to yield tetraphenylpyrazine. Since this was not observed, it is reasonable to assume that the reaction of formic acid with desylamine to form *N*-desylformamide, the glyoxaline intermediate, occurs more readily than the autocondensation of desylamine. A study of the kinetics of these reactions should be of aid in resolving these issues.

We must further examine the reason for Leuckart's quantitative yields of tetraphenylpyrazine, when Novelli, presumably employing the intermediate formamide, could get only 10 per cent. This apparent discrepancy becomes clear when it is remembered that Leuckart conducted his experiment at a temperature of 230°C., when any formamide resulting from the ammonium formate would have dissociated to yield ammonia. Thus, Novelli's, Davidson's, and Leuckart's conditions are not at all comparable, and the mechanism for each is apparently different.

TABLE 4

Pyrazines prepared by the replacement of the hydroxyl in a hydroxycarbonyl with an amino group

REACTANTS	PYRAZINE DERIVATIVE	REFERENCES
Benzoin	Tetraphenylpyrazine	(58, 74, 114, 134, 137, 162, 189, 202, 203, 204, 209)
Anisoin	Tetra-*p*-methoxyphenylpyrazine	(162)
Benzanisoin	Diphenyl-di-*p*-methoxyphenylpyrazine	(162)
p-Toluoin	Tetra-*p*-tolylpyrazine	(162)
β-Hydroxylevulinic acid	Tetramethylpyrazine	(237)
Fructose	Fructosazine	(141, 142)
Glucose	Fructosazine	(141, 142)
α-Keto-β-acetyl(or benzoyl)anilino-α,β-diphenylethane	Tetraphenylpyrazine	(79)
Benzil, benzylamine	Tetraphenylpyrazine	(111)

In table 4 we have listed not only those pyrazines whose preparation directly involves the replacement of the hydroxyl in a hydroxycarbonyl compound with an amino group, but also those for which this mechanism operates only indirectly. Thus, when benzil is treated with benzylamine under the influence of zinc chloride, tetraphenylpyrazine is obtained as a result of the intermediate formation of benzoin and ammonia.

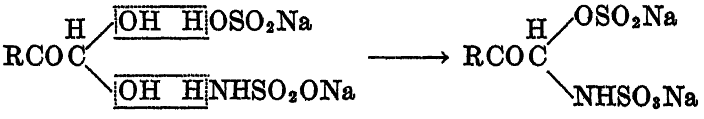
6. Treatment of the bisulfite compounds of oximino ketones with potassium cyanide

This synthesis was originated by Gastaldi (90). When an oximino ketone is treated with a strong solution of sodium bisulfite saturated with sulfur dioxide, there are apparently formed the short-lived intermediates, sulfamic acid and a hydrated glyoxal:

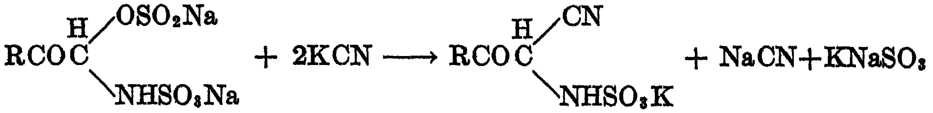
$$RCOCH{=}NOH + H_2SO_3 + H_2O \rightarrow RCOCH(OH)_2 + H_2NSO_3H$$

R = aryl or alkyl (specifically phenyl or methyl).

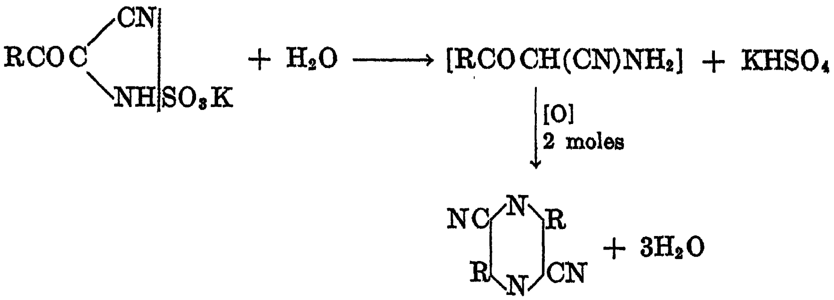
The hydrated glyoxal immediately reacts with the bisulfite and sulfamic acid (or sodium sulfamate) present, yielding a sulfite sulfamate:



Addition of a strong aqueous solution of potassium cyanide results in the replacement of the sulfite group with a cyano group:

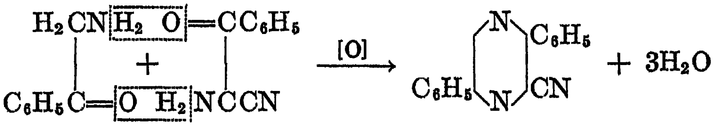


Hydrolytic cleavage at the nitrogen–sulfur bond of the cyano sulfamate yields potassium acid sulfate and the cyanoamino ketone. Finally, the latter cyclizes and undergoes oxidation to the pyrazine:



It should be noted that Gastaldi states that potassium acid sulfite results from the nitrogen–sulfur cleavage. This clearly cannot be the case if the process is only hydrolysis, as it undoubtedly is. He makes a somewhat similar statement concerning the hydrolysis of sodium sulfamate, maintaining that the products are ammonia and sodium bisulfite. Actually, they are ammonia and sodium bisulfate.

In this manner, 2,5-dimethyl-3,6-dicyanopyrazine and 2,5-diphenyl-3,6-dicyanopyrazine were synthesized. In the latter preparation some 2,5-diphenyl-3-cyanopyrazine was also isolated. This is explained by the formation of some ω-aminoacetophenone, possibly by reduction of the starting oximino compound with sulfite. This can then condense with the more normal cyanoamino ketone present in the reaction mixture:

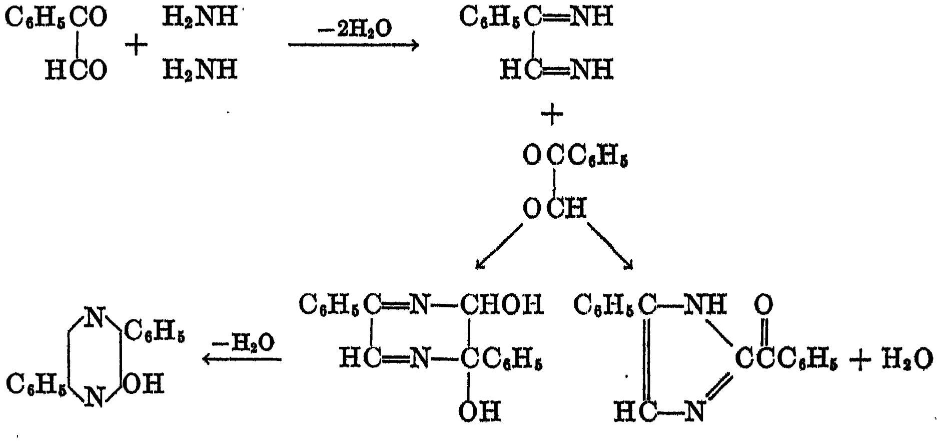


The formation of some 2,5-diphenylpyrazine by the condensation of 2 moles of ω-aminoacetophenone might also be anticipated. This was probably not detected because of the extremely limited quantity in which it would be formed.

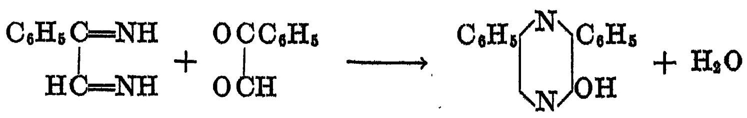
7. *Replacement of one carbonyl group in a dicarbonyl compound with an amino group*

The reaction between phenylglyoxal and ammonia was first studied by Müller and von Pechmann (156), and later was more thoroughly investigated by Pinner (175). He pointed out that one of the products isolated by the earlier workers was 2,5-diphenyl-3-hydroxypyrazine and showed that benzoylphenylglyoxaline was produced as well.

According to Pinner an intermediate diimine was first produced, which then condensed with some of the diketone present in one of two ways:



However, if this mechanism be true it is difficult to see why some 2,6-diphenyl-3-hydroxypyrazine should not form by the alternative condensation:

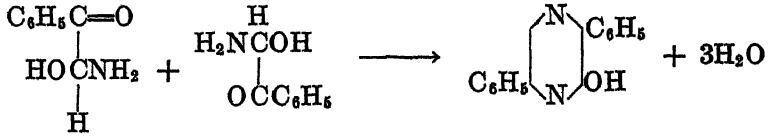


Furthermore, it was never possible to isolate any such intermediate diimine.

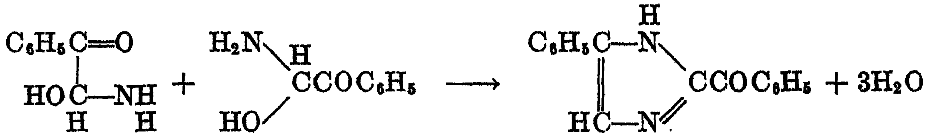
It would seem at least equally, if not more, probable that an aldehyde ammonia would first form:

$$C_6H_5COCHO + NH_3 \rightarrow C_6H_5COCHOHNH_2$$

followed by autocondensation and dehydration:



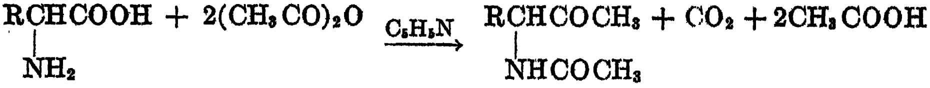
Glyoxaline formation may be accounted for as follows:



For completeness, we mention another reaction that may possibly fall in this category: that between benzil and potassium amide in the presence of ammonium chloride, from which traces of tetraphenylpyrazine have been isolated (136).

8. *The action of acetic anhydride and pyridine on amino acids*

When an α-primary amino acid containing an unsubstituted hydrogen on the α-carbon is treated with acetic anhydride in the presence of pyridine, an acetylaminoacetone is formed (54, 138):



$$\underset{\displaystyle NH_2}{RCHCOOH} + 2(CH_3CO)_2O \xrightarrow{C_5H_5N} \underset{\displaystyle NHCOCH_3}{RCHCOCH_3} + CO_2 + 2CH_3COOH$$

R = hydrogen, alkyl, or aryl.

Upon hydrolysis of the acetylaminoacetone and subsequent treatment with mercuric chloride in excess alkali, the pyrazine results:

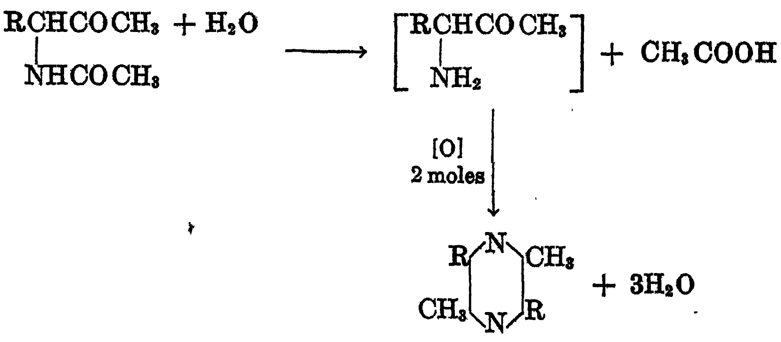


TABLE 5

Pyrazines prepared by the action of acetic anhydride and pyridine on amino acids

AMINO ACID	PYRAZINE DERIVATIVE	REFERENCES
Glycine	2,5-Dimethylpyrazine	(54)
Hippuric acid	2,5-Dimethylpyrazine	(54)
Aspartic acid	Tetramethylpyrazine	(54)
Glutamic acid	2,5-Dimethylpyrazine-3,6-dipropionic acid	(54)
Phenylalanine	2,5-Dibenzyl-3,6-dimethylpyrazine	(54, 138)
Tyrosine	2,5-*p*-Dihydroxybenzyl-3,6-dimethylpyrazine	(54, 138)
Phenylaminoacetic acid	2,5-Diphenyl-3,6-dimethylpyrazine	(138)

9. *Rearrangement of certain oximes*

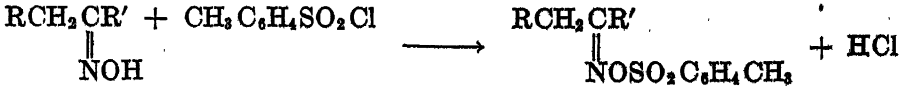
In the course of their researches on the Beckmann rearrangement, Neber and his associates (157, 158, 159, 160) discovered an interesting rearrangement of oximes (see table 6), that resulted in the formation of amino ketones.

TABLE 6

Pyrazines and dihydropyrazines prepared by the rearrangement of certain oximes

OXIME	PYRAZINE DERIVATIVE	REFERENCES
Bibenzyl ketoxime	2,5-Diphenyl-3,6-dibenzyldihydropyrazine	(160)
Benzylacetone oxime	3,6-Dibenzyl-2,5-dimethyl-3,6-dihydropyrazine	(160)
o-Nitrobenzyl methyl ketoxime	2,5-Dimethyl-3,6-di-(*o*-nitrophenyl)-dihydropyrazine	(159)
Benzyl methyl ketoxime	2,5-Dimethyl-3,6-diphenylpyrazine	(159)
Acetoacetanilide oxime	2,5-Dimethyl-3,6-(dicarboxanilide)-dihydropyrazine	(157)
2,4-Dinitrobenzyl methyl ketoxime	2,5-Dimethyl-3,6-di(2,4-dinitrophenyl)-dihydropyrazine	(158)

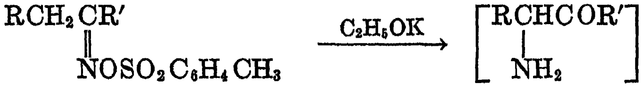
The oxime is first converted into the *p*-toluenesulfonic ester.



$$RCH_2\underset{\underset{\displaystyle NOH}{\|}}{C}R' + CH_3C_6H_4SO_2Cl \longrightarrow RCH_2\underset{\underset{\displaystyle NOSO_2C_6H_4CH_3}{\|}}{C}R' + HCl$$

R, R′ = aryl or alkyl.

Upon saponification with potassium ethoxide, rearrangement occurs with the formation of the amino ketone:



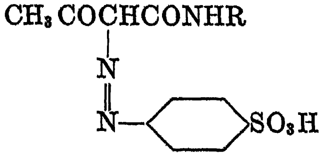
$$\underset{\overset{\|}{NOSO_2C_6H_4CH_3}}{RCH_2CR'} \xrightarrow{C_2H_5OK} \left[\underset{\overset{|}{NH_2}}{RCHCOR'}\right]$$

In most cases the amino ketones condense without further treatment into the dihydropyrazines. Although these may be oxidized to the pyrazines, this step was not carried out.

The method is claimed to be quite superior for the preparation of certain amino ketones, but there have been a number of instances where it has met with failure. Thus, negative results were reported with 3,4-dihydroxy-, 3,4-diacetoxy-, 3,4-dibenzoyloxy-, and 3,4-dimethoxyacetophenone oximes. In other cases certain modifications were found necessary. For example, the oxime of ethyl acetoacetate gave rise to methylisoxazolone, but the desired product was obtained when the anilide rather than the ester was used.

10. Reduction of diazoacetoacetanilides

Fierz-David and Ziegler (83) discovered this reaction during an investigation of certain dyestuffs. They prepared acetoacetanilides from ethyl acetoacetate and various aromatic amines. When coupled with diazotized sulfanilic acid, dyes of the form



R = aryl.

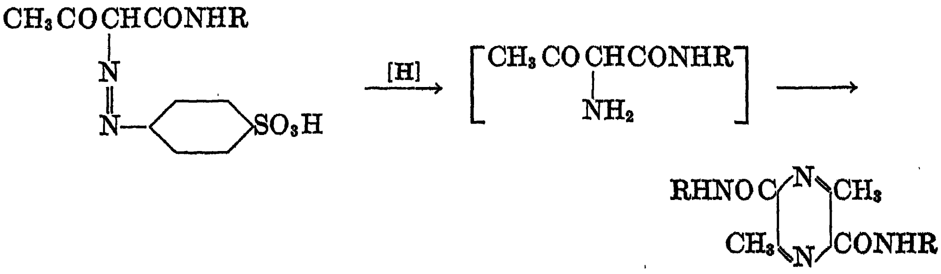
were obtained.

After reduction with stannous chloride in hydrochloric acid, the azo bond cleaved to yield the corresponding anilides of aminoacetoacetic acid. As usual, cyclization to the dihydropyrazines took place when the free bases were released with excess alkali (see table 7).

TABLE 7

Dihydropyrazines resulting from the reduction of diazoacetoanilides (83)

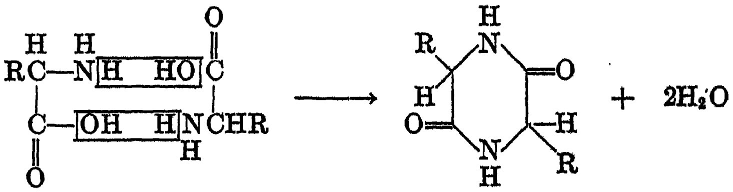
ANILIDE	DIHYDROPYRAZINE
Anilide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-dicarboxanilide
o-Toluide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-*o*-toluide
p-Toluide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-*p*-toluide
m-Xylide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-*m*-xylide
o-Anisidide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-*o*-anisidide
o-Chloroanilide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-*o*-chloroanilide
2,5-Dichloroanilide	2,5-Dimethyl-3,6-dihydropyrazine-3,6-2,5-dichloroanilide



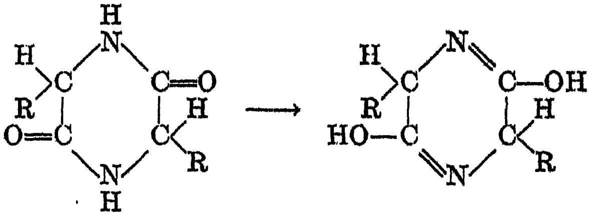
These workers did not attempt to oxidize the dihydropyrazines to the pyrazines.

11. *Enolization of diketopiperazines*

Although the diketopiperazines (or dipeptide anhydrides) which may be formed by the condensation of 2 moles of amino acid (generally as the ester)

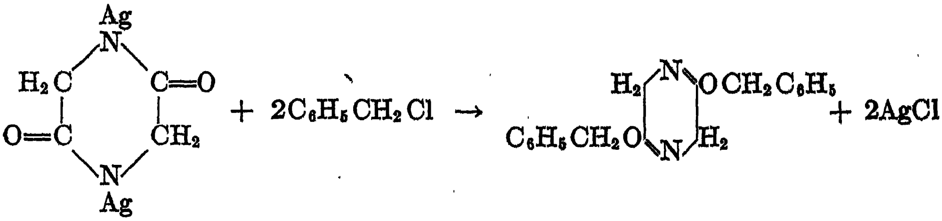


are not directly within our scope, they nevertheless enter the discussion, since it was observed that they may be made to enolize into the dihydropyrazines:



Other enolization products are possible, but experimental evidence favors this particular species. The investigators have not reported any attempts to oxidize these compounds to the pyrazine derivatives.

Two methods have been used to induce enolization. Karrer and his coworkers (120, 121) refluxed the silver salt of glycine anhydride with benzyl chloride to obtain 2,5-dibenzyloxydihydropyrazine.



When *p*-carbethoxybenzyl chloride was employed as the halide, the corresponding 2,5-di(*p*-carbethoxy)benzyloxydihydropyrazine resulted (121).

Abderhalden and his associates (1, 2) simply refluxed the diketopiperazines with aniline and then removed the aniline with ether. They reported successful conversions for the anhydrides of glycine, alanine, sarcosine, and leucylglycine. In one instance (alanine), they were able to form the dihydropyrazine directly, by heating the amino acid ester with aniline.

12. Miscellaneous

We have included in this category a number of assorted reactions which apparently depend upon the intermediate formation of α-primary aminocarbonyl compounds. In many cases the production of these substances is unexpected. Without experimental verification it would be pointless to attempt any detailed interpretation of the chemistry involved. Therefore, we present only the salient information as embodied in table 8.

TABLE 8

Miscellaneous methods for preparing pyrazines

REACTANTS	CONDITIONS	PRODUCTS	REFERENCES
α,β-Dichlorostyrene and ammonia	Elevated temperature and pressure	2,5-Diphenylpyrazine	(130)
α-Methyl-α,β-dichlorostyrene and ammonia	Elevated temperature and pressure	2,5-*p*-Ditolylpyrazine	(130)
Dypnone ($C_6H_5(CH_3)C{=}CHCOC_6H_5$) and hydroxylamine	Large excess of hydroxylamine at room temperature for several days	2,5-Diphenylpyrazine	(102)
α-(4-Morpholinyl)propiophenone and ammonium carbonate	105°C.; 15 hr. under pressure	2,5-Dimethyl-3,6-diphenylpyrazine	(45)

B. AUTOCONDENSATION OF α-HYDROXYIMINES

When Minovici (153) treated mandelonitrile with dry hydrogen chloride in anhydrous ether, he obtained a yellow crystalline substance which Japp and his associates (112, 113) proved to be 3-keto-2,5-diphenyl-3,4-dihydropyrazine (or its tautomer, 2,5-diphenyl-3-hydroxypyrazine). Ingham (109) studied the reaction more closely and provided the following scheme for its mechanism.

As would be expected from the conditions an iminochloride first forms:

$$HCl + RCH(OH)CN \xrightarrow{\text{ether}} RCH(OH)CCl{=}NH$$

Cyclization then occurs by the loss of two molecules of water:

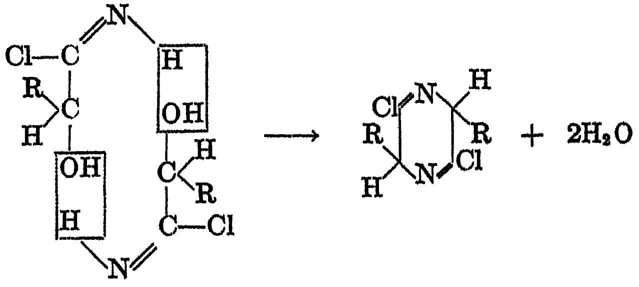
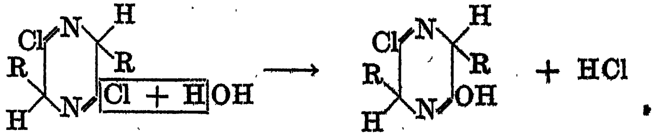


TABLE 9

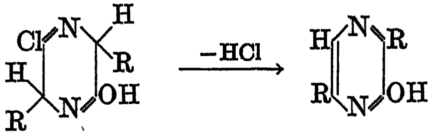
Pyrazines prepared by the autocondensation of α-hydroxyimines

REACTANTS	PYRAZINE DERIVATIVE	REFERENCES
Mandelonitrile and α-naphthylmagnesium bromide	2,5-Diphenyl-3,6-di-α-naphthylpyrazine	(150)
Benzoin oxime and hydrogen	Tetraphenylpyrazine	(229)
Mandelonitrile	3-Keto-2,5-diphenyl-3,4-dihydropyrazine	(112, 113, 153)
Anisaldehyde cyanohydrin	3-Keto-2,5-di-*p*-methoxyphenyl-3,4-dihydropyrazine	(148)
Cinnamaldehyde cyanohydrin	3-Keto-2,5-distyryl-3,4-dihydropyrazine	(148)
m-Nitrobenzaldehyde cyanohydrin	3-Keto-2,5-di-*m*-nitrophenyl-3,4-dihydropyrazine	(109)
p-Nitrobenzaldehyde cyanohydrin	3-Keto-2,5-di-*p*-nitrophenyl-3,4-dihydropyrazine	(109)

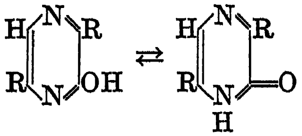
Interaction with one molecule of water leads to the elimination of a molecule of hydrogen chloride:



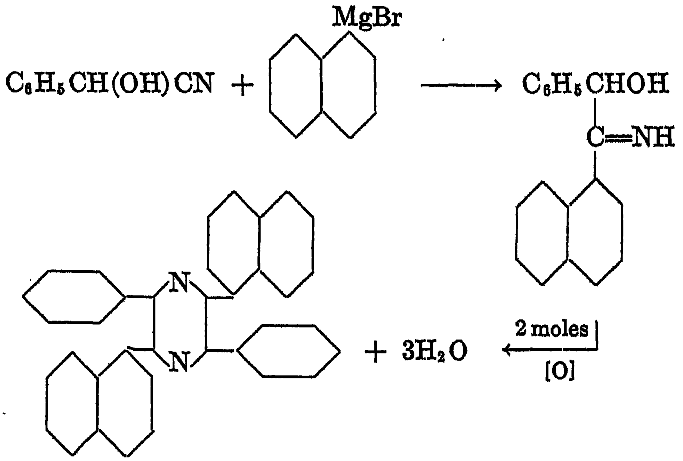
Finally, another mole of hydrogen chloride is lost to yield the product:



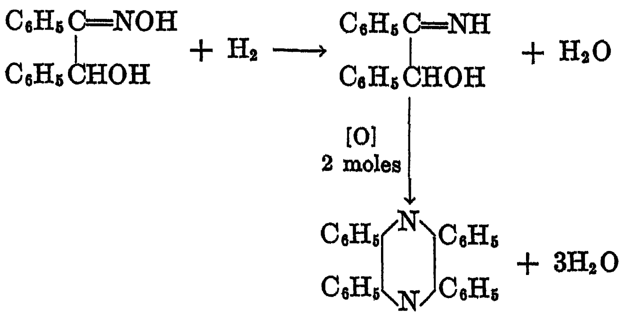
The hydroxypyrazine is a tautomer of the keto dihydropyrazine, as Japp an Knox (112) have shown:



McKenzie and Kelman (150) modified the procedure by treating the aldehyd cyanohydrin with the Grignard reagent instead of hydrogen chloride. The thus prepared 2,5-diphenyl-3,6-di[α-naphthyl]pyrazine from mandelonitrile an α-naphthylmagnesium bromide.



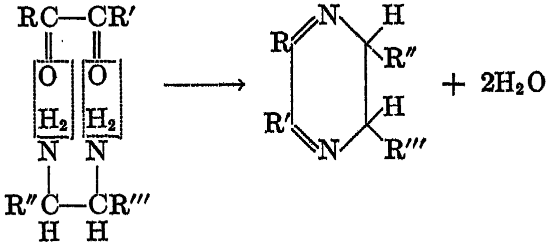
A similar cyclization mechanism is ostensibly in operation in the hydrogenation of benzoin oxime (229):



C. CONDENSATION OF α,β-DIAMINES WITH α,β-DICARBONYLS

1. The direct formation of pyrazines

This method of forming the pyrazine ring has proved of wide utility. The general procedure is to react the dicarbonyl with the diamine in a mutual solvent.



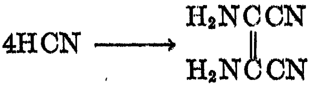
R,R′,R″,R‴ = aryl, alkyl, or hydrogen.

The resulting 2,3-dihydropyrazines may be isolated, or oxidized directly to the pyrazines (see table 10).

If R and R′ are aryl, the reaction proceeds smoothly and the yields are uniformly good. Again, difficulties are encountered (15) when attempts are made to prepare the first member of the series by the condensation of glyoxal with ethylenediamine (118). Apparently with those low-molecular-weight substances, the conditions as ordinarily employed favor the formation of polymers, e.g.:

$$=NCH_2CH_2N=CHCH=NCH_2-$$

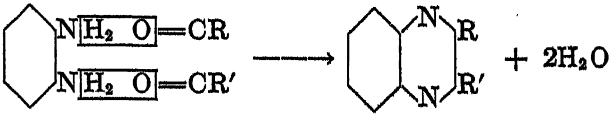
A novel method for the preparation of 2,3-dicyanopyrazines depends on the use of the tetramer of hydrogen cyanide, diaminomaleonitrile, as the diamine (96, 140):



$$4HCN \longrightarrow \begin{array}{c} H_2NCCN \\ \| \\ H_2NCCN \end{array}$$

2. The oxidation of quinoxaline or its derivatives

The value of the method may be extended by means of quinoxaline and its derivatives prepared by the condensation of *o*-phenylenediamine with various dicarbonyl compounds.



These compounds are important, since they may be easily oxidized with dilute alkaline permanganate to pyrazinedicarboxylic acids in good yields (see table 11):

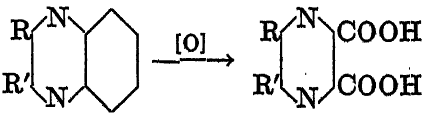


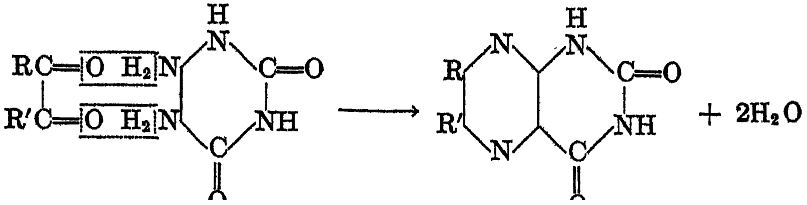
TABLE 10

Pyrazines prepared by the condensation α,β-diamines with α,β-dicarbonyls

REACTANTS	PYRAZINE DERIVATIVE	REFERENCES
Diaminomaleonitrile and glyoxal	2,3-Dicyanopyrazine	(96, 140)
Diaminomaleonitrile, glyoxal, and hot water	6-Hydroxy-2,3-dicyanopyrazine	(104)
Diacetyl and ethylenediamine	2,3-Dimethylpyrazine	(118)
Benzil and ethylenediamine	2,3-Diphenylpyrazine; also some tetraphenylpyrazine	(5, 145, 146, 147)
Benzil and propylenediamine	2,3-Diphenyl-5-methylpyrazine	(145, 146, 147)
Anisil and ethylenediamine	2,3-Dimethoxyphenylpyrazine; also some tetra[dimethoxyphenylpyrazine]	(146, 147)
Methyl propyl diketone and ethylenediamine	2-Methyl-3-propylpyrazine	(31)
Stilbenediamine and benzil	Tetraphenylpyrazine	(81)
Diaminomaleonitrile and diacetyl	5,6-Dimethyl-2,3-dicyanopyrazine	(96, 140)
Diaminomaleonitrile and benzil	5,6-Diphenyl-2,3-dicyanopyrazine	(96, 140)
Phenylethylenediamine and benzil	2,3,5-Triphenylpyrazine	(81)
α-Dicyanodibenzyl diketone and ethylenediamine	2,3-Di[α-cyanobenzyl]-5,6-dihydropyrazine	(67)
3-Chlorotrimethylene-1,2-diamino-dihydrochloride and benzil	5,6-Diphenyl-2-chloromethyl-2,3-dihydropyrazine	(174)
1,2-Diaminocyclobutane and benzil	Tetraphenylpyrazine	(33)

3. The cleavages of lumazines

As a further application of this general reaction we may consider the formation of lumazines by the reaction of 4,5-diaminouracil with various dicarbonyl compounds:



Weijlard, Tishler, and Erickson (227) found that these derivatives could be split with rupture of the pyrimidine nucleus, the pyrazine nucleus remaining intact (see table 12).

Alkaline cleavage led to the formation of 2-amino-3-pyrazinoic acid derivatives:

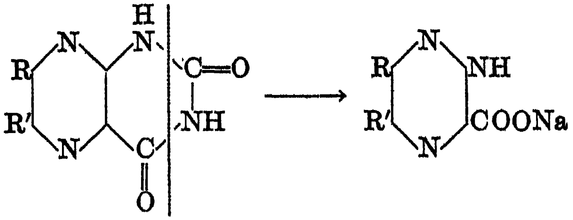
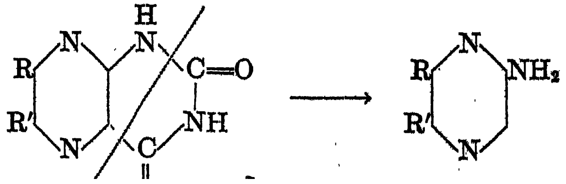


TABLE 11

Pyrazinecarboxylic acids prepared by the oxidation of quinoxalines

QUINOXALINE	PYRAZINE ACID	REFERENCES
Quinoxaline	2,3-Pyrazinedicarboxylic acid	(20, 21, 60, 73, 89, 116, 117)
α,β-Diphenyl-1,2-naphthoquinoxaline (or its quinone)	2,3-Diphenyl-5-carboxypyrazine-6-*o*-benzoic acid	(46)
β-Phenyl-1,2-naphthoquinoxaline-3,4-quinone	Pyrazine-3-phenyl-5-carboxy-6-*o*-benzoic acid	(48)
Naphtho-1,2-quinoxalyl-β-acetic acid	6-Carboxy-5-(*o*-carboxyphenyl)-2-pyrazineacetic acid	(47)
2-Methylquinoxaline	5-Methyl-2,3-pyrazinedicarboxylic acid	(22, 135)
2,3-Dimethylquinoxaline	5,6-Dimethylpyrazinedicarboxylic acid	(89)
2-Ethoxymethylquinoxaline	Pyrazinetricarboxylic acid	(23)
2,3-Quinoxalinedicarboxylic acid anhydride	Pyrazinetetracarboxylic acid	(37)

With sulfuric acid, on the other hand, further decarboxylation occurred, resulting in aminopyrazines:



D. THE ACTION OF AMMONIA OR ITS SALTS ON POLYHYDROXY COMPOUNDS

Etard (78) and Stoehr (202) discovered independently that 2,5-dimethylpyrazine (together with much smaller quantities of 2,5-dimethyl-3-ethylpyrazine) could be obtained by distilling glycerol with a mixture of ammonium chloride and ammonium phosphate. If only ammonium phosphate was used the distillate was found to consist primarily of pyridines, predominantly β-picoline.

The postulated mechanism for both cases involves the intermediate formation of acrolein:

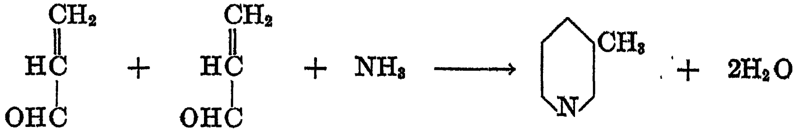
$$CH_2OHCHOHCH_2OH \rightarrow CH_2{=}CHCHO + 2H_2O$$

TABLE 12

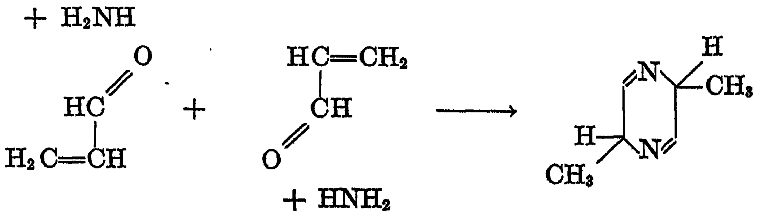
Pyrazine derivatives prepared by Weijlard, Tishler, and Erickson (227)

R	R′	CLEAVAGE MEDIUM	PRODUCT
H	H	NaOH	2-Amino-3-pyrazinoic acid (2-hydroxy-3-pyrazinoic acid under more stringent conditions)
CH_3	H	NaOH	2-Amino-6-methyl-3-pyrazinoic acid
C_6H_5	C_6H_5	NaOH	2-Amino-5,6-diphenyl-3-pyrazinoic acid
CH_3	CH_3	NaOH	2-Amino-5,6-dimethyl-3-pyrazinoic acid
H	H	H_2SO_4	Aminopyrazine
CH_3	H	H_2SO_4	2-Amino-6-methylpyrazine
C_6H_5	C_6H_5	H_2SO_4	2-Amino-5,6-diphenylpyrazine
CH_3	CH_3	H_2SO_4	2-Amino-5,6-dimethylpyrazine
C_6H_5	H	H_2SO_4	2-Amino-5(or 6)phenylpyrazine

Two molecules of acrolein may combine with one of ammonia to form β-picoline:



On the other hand, if 2 molecules of acrolein combine with 2 molecules of ammonia, pyrazines result:



The 2,5-dimethyl-2,5-dihydropyrazine is further oxidized under the reaction conditions to yield finally 2,5-dimethylpyrazine. A possible explanation for the occurrence of 2,5-dimethyl-3-ethylpyrazine may be found in Stoehr's observation that the addition of acetaldehyde increases the yield of this compound (209).

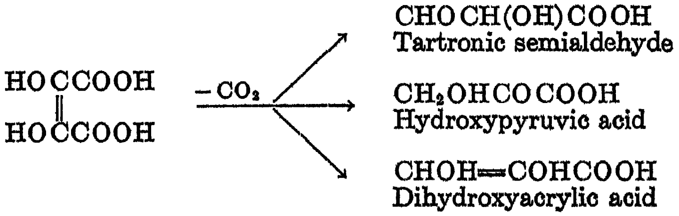
Baeyer (10) has obtained β-picoline from the dry distillation of acrolein ammonia; however, support of the mechanism for the production of dimethylpyrazine is lacking.

The yield of dimethylpyrazine is poor (5–10 per cent). Attempts to improve it by variation of conditions and the use of catalysts such as charcoal and copper sulfate have not been substantially effective (12, 62, 78, 97, 98, 129, 202, 203, 204).

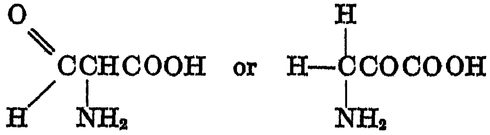
The method has been modified by employing polyhydroxy compounds other than glycerol. Dori and Mohring (65) have patented a procedure requiring the treatment of sucrose with ammonia or ammonium salts at 200°C. and 50 atm. in the presence of a sulfite, bisulfite, animal charcoal, or a copper sulfate catalyst. The process was claimed to yield "basic mixtures including pyrazines."

A similar reaction appears to be that of ammonium hydroxide and glucose at 100°C. (25, 215, 217). Amongst the products identified were pyrazine, methylpyrazine, and 2,6-dimethylpyrazine. Since pyrazines other than 2,5-dimethylpyrazine were obtained, it would seem to indicate that mechanisms other than that postulated for glycerol might also be involved.

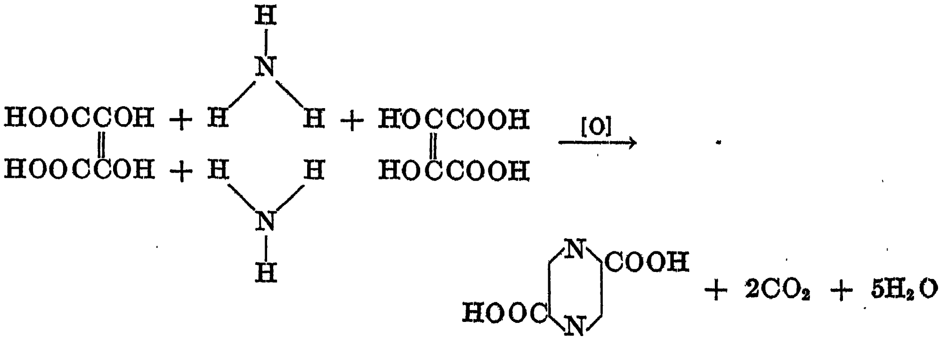
Fenton (82) discovered that when ammonium dihydroxymaleate was allowed to remain in contact with aqueous ammonia at 50–60°C. for half an hour, 2,5-pyrazinedicarboxylic acid was formed. He pointed out that there were several routes by which such a change could take place. Thus, by the loss of carbon dioxide, dihydroxymaleic acid may give one of several products:



Condensation of any of these with ammonia could lead to an α-primary aminocarbonyl carboxylic acid such as



Cyclization and oxidation would then yield the 2,5-pyrazinedicarboxylic acid. Fenton indicates another possibility in the formation of a tetracarboxylic acid which could give the product by oxidation and loss of carbon dioxide:



This, however, seems more remote, since we know that the pyrazinecarboxylic acids require much more drastic conditions than those present in order to undergo decarboxylation.

It is noteworthy that when the esters were used instead of the free acids, oxamide was formed with no traces of any pyrazines.

E. CERTAIN REACTIONS OF THE PHENACYLBENZYLAMINES

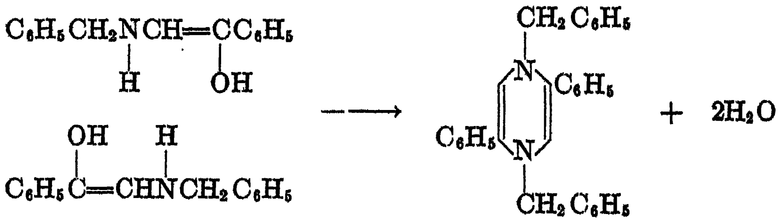
Mason and his students (147) carried out a series of studies on the preparation and properties of the phenacylbenzylamines, which led to the synthesis of a number of pyrazine derivatives.

The mono- and diphenacylamines may be prepared by heating a mixture of benzylamine and ω-bromoacetophenone:

$$2C_6H_5CH_2NH_2 + 3C_6H_5COCH_2Br \longrightarrow$$

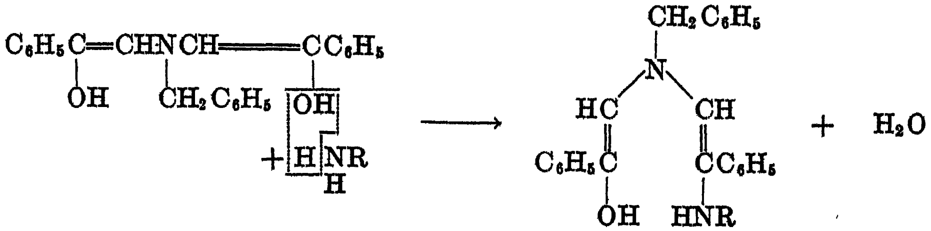
$$\underset{\text{Monophenacylbenzylamine}}{C_6H_5COCH_2NHCH_2C_6H_5 \cdot HBr} + \underset{\text{Diphenacylbenzylamine}}{(C_6H_5COCH_2)_2NCH_2C_6H_5 \cdot HBr} + HBr$$

When monophenacylbenzylamine hydrobromide is treated with potassium hydroxide at 100°C., 1,4-dibenzyl-2,5-diphenylpyrazine dihydride is formed:



Boiling this product causes debenzylation, resulting in 2,5-diphenylpyrazine and 2 molecules of toluene.

The formation of pyrazine from diphenacylbenzylamine occurs in a somewhat different manner. The latter in its enolic form may react with an amine:

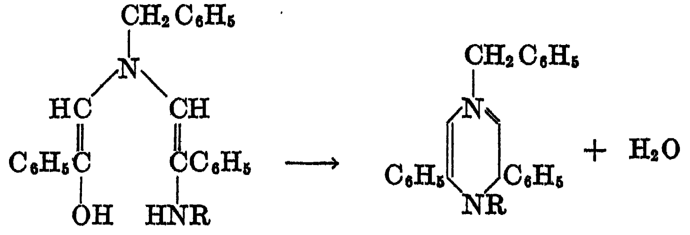


R = aryl or hydrogen.

Aminodiphenacylbenzylamine

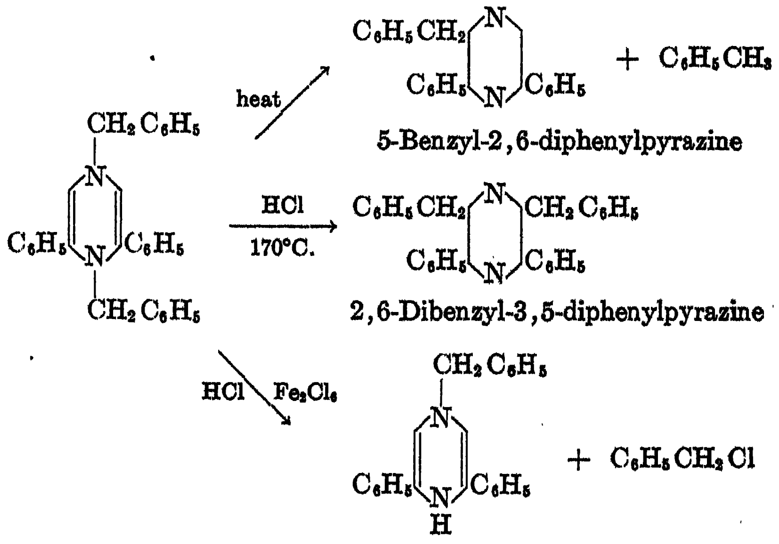
Such an intermediate was actually isolated when the amine reagent was cold ammonia (R = H).

With heat, another molecule of water is eliminated, effecting cyclization:



Compounds thus prepared were 1-benzyl-3,4,5-triphenyldihydropyrazine (R = phenyl); 1,4-dibenzyl-3,5-diphenyldihydropyrazine (R = benzyl); and 1-benzyl-3,5-diphenyldihydropyrazine (R = H). 1-Benzyl-3,5-diphenyldihydropyrazine may be converted by heat to 3,5-diphenylpyrazine with the loss of a molecule of toluene (see, also, reference 172).

Depending on the reagent, 1,4-dibenzyl-3,5-diphenyldihydropyrazine may undergo debenzylation and/or rearrangement:



5-Benzyl-2,6-diphenylpyrazine

2,6-Dibenzyl-3,5-diphenylpyrazine

1-Benzyl-3,5-diphenyldihydropyrazine

F. THE ACTION OF HYDRAZINE OR HYDROXYLAMINE ON DIHYDROXYMORPHOLINE

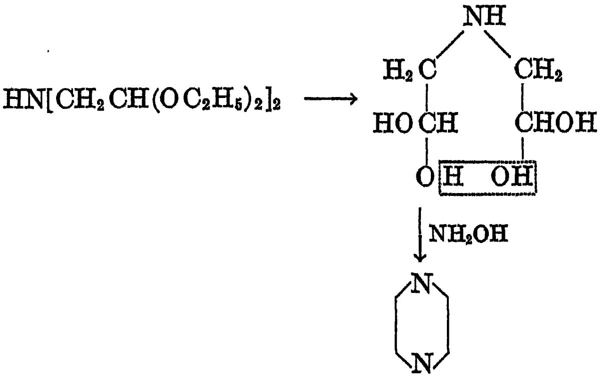
Wolff and Marburg (247) discovered that the heterocyclic oxygen atom in a morpholine could be replaced with nitrogen, thus affording a means of transition to the pyrazine series.

They first prepared diacetalylamine by heating chloroacetal with ammonia in a sealed tube:

$$2ClCH_2CH(OC_2H_5)_2 + NH_3 \rightarrow HN[CH_2CH(OC_2H_5)_2]_2 + 2HCl$$

Conversion to pyrazine was then accomplished by heating this substance with the hydrochloride of either hydrazine or hydroxylamine under pressure.

The dihydroxymorpholine which is formed (and which may be isolated) is thereby transformed into pyrazine:



Although inconvenient because of the long periods of heating in closed vessels in both steps, this method remains the best of the published pyrazine syntheses with regard to yield—78 per cent based on diacetalylamine (107).

G. THE OXIDATION OF PIPERAZINES

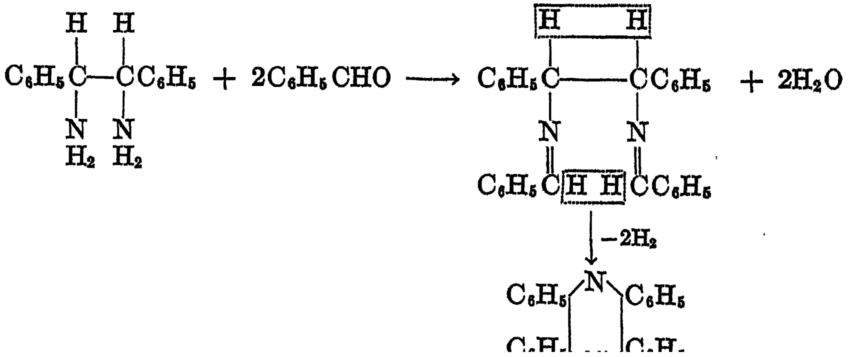
In order to oxidize piperazine to pyrazine, Stoehr (207) found it necessary to distill either the base or its hydrochloride over zinc dust. He attributed the oxidizing action of the zinc to the fact that it contained zinc oxide, and indeed found that the best results were obtained by using a mixture of zinc dust and lime. As might be expected the yields were very poor, not exceeding 10 per cent.

Recently, Dixon (64) obtained a patent based on the vapor-phase dehydrogenation of piperazine. The piperazine in the form of a dilute solution in benzene was passed in the vapor phase over copper chromite at 215–500°C. Yields of 35–40 per cent pyrazine were reported.

H. THE ACTION OF STILBENEDIAMINE ON BENZALDEHYDE

Grossman (95) showed that the primary product of the reaction between stilbenediamine and benzaldehyde is dibenzylidenestilbenediamine.

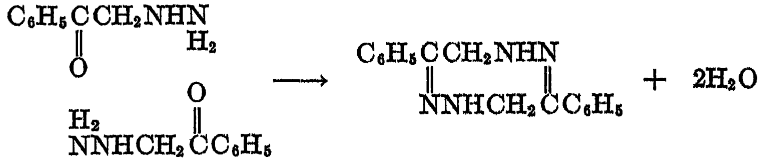
If the reaction is forced, dehydrogenation may occur, forming tetraphenylpyrazine:



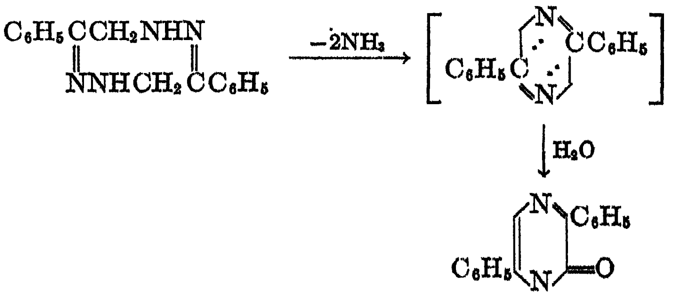
To obtain the final product from the initial reactants, it was found necessary to heat the mixture in ethanol at 180°C. for 10 hr. in a sealed tube.

I. THE CONDENSATION OF PHENACYLHYDRAZINES

Phenacylhydrazine apparently possesses the property of autocondensation, in a manner resembling that of the α-primary aminocarbonyl compounds, to form an eight-membered heterocycle (35).



This is followed by a spontaneous ring contraction with elimination of two molecules of ammonia and the addition of one of water:



p-Bromophenacylhydrazine after a similar course of reactions yields 2,5-di-*p*-bromophenyl-3-keto-3,4-dihydropyrazine.

J. BIOLOGICAL SYNTHESES

In the introduction we mentioned the condensed pyrazine derivative, riboflavin. There are also indications that the simple nucleus may take part in life processes. Scattered throughout the literature are a number of reports by various workers concerning the occurrence of such derivatives in the fusel oils resulting from the fermentation of various carbohydrate products. Table 13 presents a résumé of the data.

The exact mechanism for the production of these products is as yet unknown, but even disregarding possible degradations of riboflavin, the fact that they do occur should not occasion surprise. In the previous sections we have seen how many biologically important substances may serve in pyrazine syntheses. We may cite the action of ammonia on various polyhydroxy compounds such as glycerol and sugars, as well as the formation of pyrazines from amino acids.

In this connection we may recall an interesting experiment performed by Kikoji and Neuberg (122), in which they fed rabbits doses of aminoacetaldehyde hydrochloride. Examination of the urine revealed that cyclization had occurred *in vivo*, resulting in the formation of pyrazine which could be isolated.

Stolte (215) reported the biological oxidation of the pyrazine derivative fructosazine. This was fed to rabbits, which oxidized it to 2-methylol-5-pyrazinoic acid. The oxidation product was isolated from the urine.

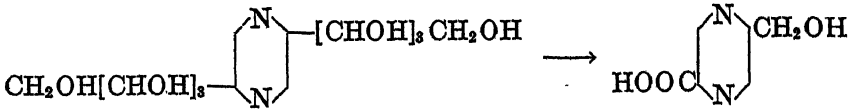


TABLE 13

Pyrazine derivatives isolated from fusel oil

PYRAZINE DERIVATIVE	CARBOHYDRATE SOURCE OF THE FUSEL OIL	REFERENCE
2,5-Dimethylpyrazine	(Kahlbaum's amyl alcohol)	(11)
Trimethylpyrazine	Molasses	(155)
Trimethylpyrazine	Potatoes	(191)
Tetramethylpyrazine	Potatoes	(191)
Diethylpyrazine	Potatoes	(191)
Triethylmethylpyrazine	Potatoes	(191)
2,5-Diethylpyrazine	Beet molasses	(219)

K. UNCLASSIFIED

There are a number of reactions found in the literature that are productive of pyrazine derivatives but for which the mechanisms are obscure. These will be found grouped in table 14.

V. PROPERTIES

A. GENERAL

The simple alkylpyrazines, like the corresponding pyrimidines and pyridazines, are low-melting solids which yield colorless mobile liquids. If not carefully purified they darken rapidly on standing. The compounds possess distinctive odors which generally have been described as narcotic.

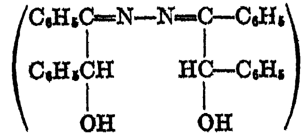
A characteristic property is the extreme solubility of the lower members in the solvents alcohol, ether, and water. These same compounds are hygroscopic and form hydrates with great ease. At least in one case the difficulty of removing the final traces of water has led to incorrect analytical results with consequent misinterpretation of structure (Gutknecht's work reported in the historical section).

Because of their volatility with steam, codistillation with steam is of great aid in the separation of these bases from reaction mixtures. Pfann (173) has reported the existence of a water–pyrazine azeotrope.

Although no quantitative measurements have been made, the behavior of the pyrazines indicates that they are weaker bases than the pyridines. This might have been anticipated from a consideration of the negative inductive effects of the nitrogen atoms in the ring, which would mutually bind each basic lone pair of electrons more tightly to the ring. Since the inductive effect is exerted more effectively in the ortho and para positions, it would be anticipated that pyridazine and pyrazine would be less basic than pyrimidine. Such a relationship holds true in the five-membered nitrogen heterocycles, where the basic dissociation constants of pyrazole and imidazole have been reported as 3×10^{-13} and 1.2×10^{-7}, respectively (188a).

TABLE 14

Unclassified methods for the preparation of pyrazines

REACTANTS	CONDITIONS	PRODUCTS	REFERENCES
Benzaldehyde and formamide	Reflux equal quantities	Tetraphenylpyrazine	(34)
Benzaldehyde and magnesium nitride	Heat under pressure	Tetraphenylpyrazine	(195)
α-Aminophenylacetonitrile $(C_6H_5CH(CN)NH_2)$ and chloral	Anhydrous ether	2,3-Dichloro-6-phenylpyrazine	(154)
α-Aminophenylacetonitrile and bromal	Anhydrous ether	2,3-Dibromo-6-phenylpyrazine	(154)
α-Amino-*p*-methoxyphenylacetonitrile and chloral	Anhydrous ether	2,3-Dichloro-6-*p*-methoxyphenylpyrazine	(154)
α-Amino-*p*-methoxyphenylacetonitrile and bromal	Anhydrous ether	2,3-Dibromo-6-*p*-methoxyphenylpyrazine	(154)
Bisbenzoinketazine  and hydrazine	Heat	Tetraphenylpyrazine	(52, 53)
Bis-*p*-toluoinketazine	Heat	Tetra-*p*-tolylpyrazine	(53)
Amarine (4,5-dihydro-2,4,5-triphenyl imidazole)	Reduction with amalgams in ethanol	Tetraphenylpyrazine (a minor product)	(216)

In contrast, the methyl-substituted compounds are more basic because of the positive inductive effect of the alkyl group. The completely reduced piperazines are much stronger bases than the parent pyrazines, which possess an aromatic system of conjugated double bonds.

Advantage is taken of the basicity for the isolation of these compounds from aqueous solutions by adding excess alkali, whereby they may be forced into another phase, despite their extreme solubility in pure water.

Notwithstanding the two basic nitrogen atoms in the ring, the early workers generally considered the pyrazines to be monoacid bases, since only one molecule of an alkyl halide was found to react with one of a pyrazine to form a monoquaternary salt. However, Tutin and Caton (226) later reported, on spectroscopic and chemical evidence, that the two series of salts could be formed. Studies based on potentiometric titrations are lacking.

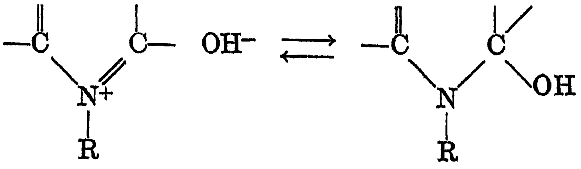
Durio and Bissi (66) have shown that both nitrogens may coördinate with oxygen in the compound which they prepared by the action of alkaline ferrocyanide on *α*-benzil dioxime, dioxotetraphenylpyrazine. This compound may be deoxygenated to tetraphenylpyrazine and some tetraphenylpiperazine with zinc in glacial acetic acid. Cripps and Perroncito (48) presumably were able to prepare a derivative in which only one of the nitrogens was oxygenated by treating 2,3-diphenyl-5-carboxypyrazine-6-*o*-benzoic acid with hydrogen peroxide in glacial acetic acid.

The complex salts formed with mercuric chloride are of importance. These are white insoluble products which may be easily decomposed by acid or alkali to liberate the bases, so that they may serve as a means for the isolation and purification of these compounds.

For detailed methods of the preparation and properties of many of the salts of the simple pyrazines the reader is referred to the various papers of Stoehr and Wolff (bibliography) in addition to that of Tutin and Caton previously cited.

B. THE QUATERNARY SALTS

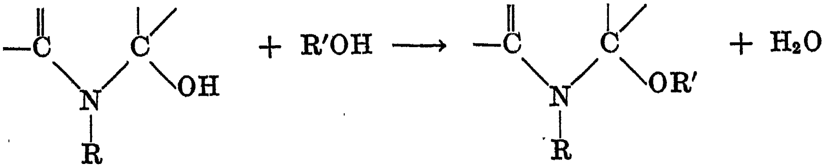
Aston carried out a series of investigations of the quaternary salts designed to give a more thorough insight into the structures of these compounds (7, 8). It is well known in the pyridine and quinoline series (192) that a quaternary hydroxide may undergo one of two possible changes, depending upon the nuclear substituents. There may be formed a pseudo base (II):



I
True base

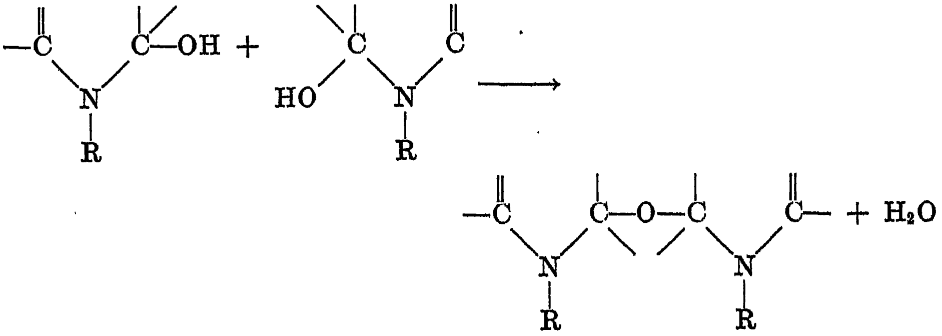
II
Pseudo base

As is shown, an equilibrium may exist between the electrolyte (I) and the nonelectrolyte (II). There is a further difference between I and II in that the latter possesses one less double bond than the former. The pseudo base may then react with an alcohol to give an ether (III):



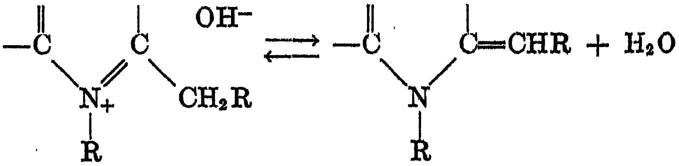
III
Ether

And in some cases another molecule of the pseudo base may function as the alcohol, yielding the dimeric ether (IV):



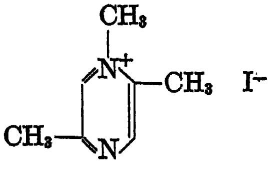
IV
Dimeric ether

However, if the substituent on the carbon carrying the hydroxyl of the pseudo base is not hydrogen and is attached by a methylene group, an anhydro base (V) may also be formed by the elimination of water:

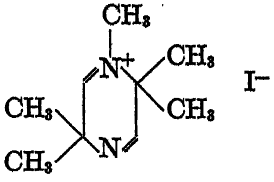


I V
Anhydro base

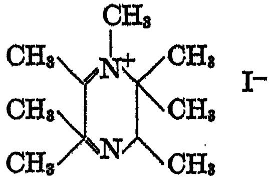
Aston worked with three quaternary salts of the pyrazine series: 1,2,5-trimethylpyrazinium iodide (A), 1,2,2,5,5-pentamethyldihydropyrazinium iodide (B), and 1,2,2,3,5,5,6-heptamethyldihydropyrazinium iodide (C):



A



B

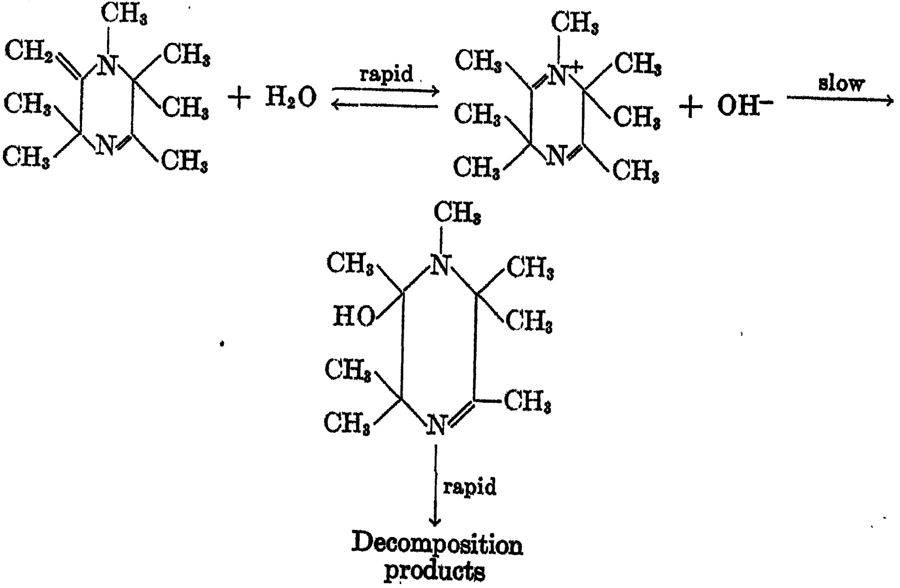


C

Upon the addition of potassium hydroxide to these iodides, B was found to yield the pseudo base (II), and C the anhydro base (V), but no base could be isolated from A. The reaction rates and basic dissociation constants of the equilibrium mixtures were determined from conductance measurements in aqueous solution. It was shown that for B an equilibrium between forms I and II is reached extremely rapidly, and is in favor of II.

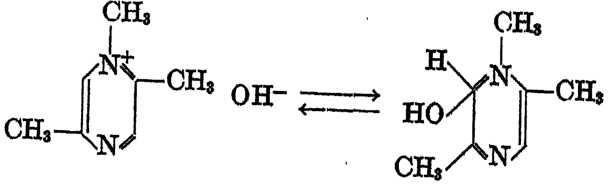
In the case of C an equilibrium between I and V is also attained very rapidly. Although the solution was strongly basic ($K = 4 \times 10^{-3}$) when freshly prepared, the basicity diminished in time simultaneously with the formation of decomposition products. Furthermore, a kinetics study revealed that the decomposition rate was proportional to the first power of the concentration of the pyrazinium and hydroxide ions. Consideration of these facts led to the conclusion that the formation of the pseudo base is the controlling step.

These transformations may be indicated as follows:



The failure of the attempts to isolate a base from A was further probed kinetically. Experiments showed that the disappearance of strong base from a mixture of 1,2,5-trimethylpyrazinium iodide and sodium hydroxide proceeds in two main directions.

In both cases there first occurs a rapid equilibrium favoring the left:

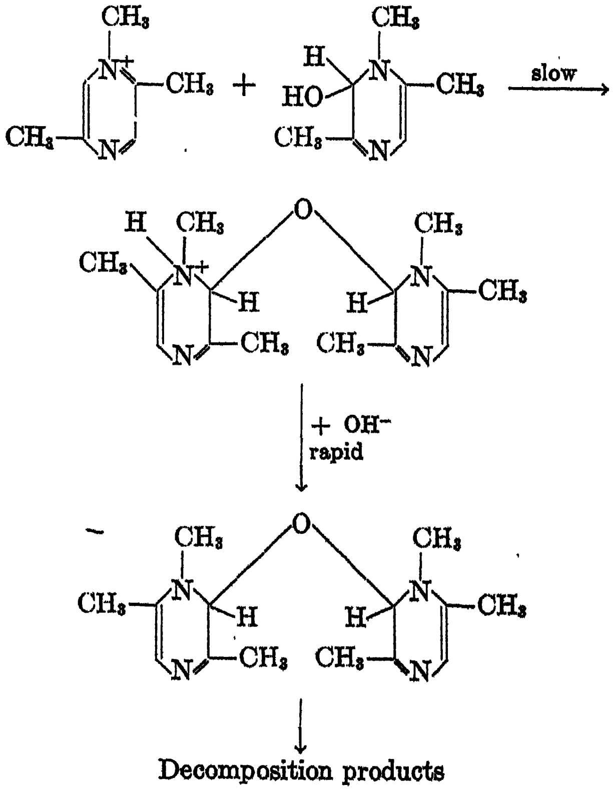
 (1)

(Py+) (POH)

The equilibrium constant, K_1, is given by:

$$K_1 = \frac{[\mathrm{POH}]}{[\mathrm{Py}^+][\mathrm{OH}^-]}$$

Consider now the formation of a dimeric ether in accordance with the following mechanism:



(2)

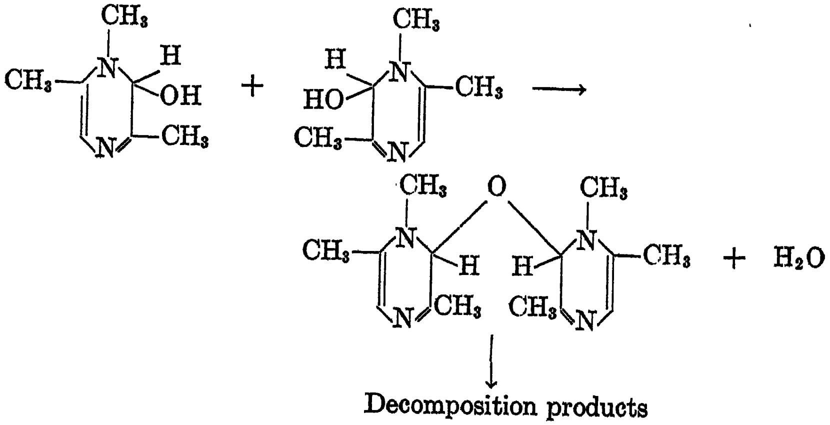
Reaction 2 is thus the rate-controlling step. The differential equation is:

$$-\frac{d[\mathrm{Py}^+]}{dt} = k_2[\mathrm{POH}][\mathrm{Py}^+]$$

Substituting for [POH] from the expression for K_1, there results:

$$-\frac{d[\mathrm{Py}^+]}{dt} = \mathrm{K}_1 k_2[\mathrm{Py}^+]^2[\mathrm{OH}^-] = k'[\mathrm{Py}^+]^2[\mathrm{OH}^-]$$

On the other hand, if the slow reaction takes place by the direct condensation of two molecules of the pseudo base:



(3)

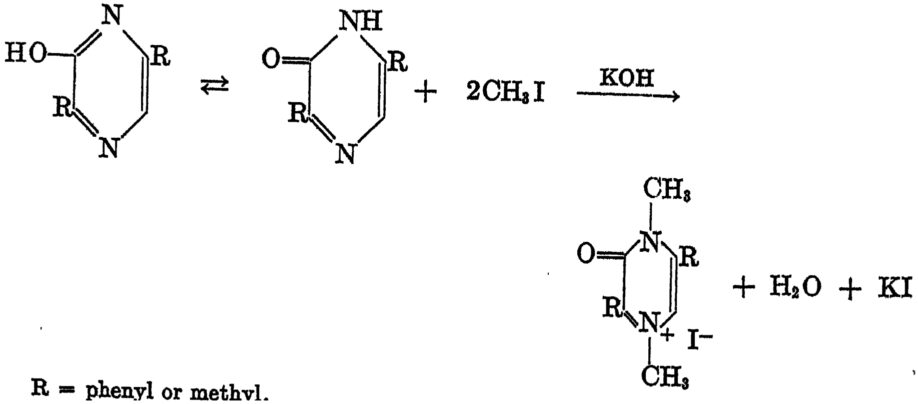
the kinetics will be represented by:

$$-\frac{d[Py^+]}{dt} = -\frac{d[POH]}{dt} = k_3[POH]^2 = K_1^2k_3[Py^+]^2[OH^-]^2$$
$$= k''[Py^+]^2[OH^-]^2$$

These two mechanisms are supported by the data, the latter one predominating in solutions which contain an excess of hydroxide ion.

These results were interpreted by Aston in the light of ring conjugation. Since the —N=C— double bond of B will be more prone to add the hydroxyl ion than that of the fully conjugated system A, the I–II equilibrium will be displaced more to the side of II in the former case. Analogously, the non-conjugation of C leads to the rapid establishment of the I–V equilibrium, and again the fact that this does not occur in the case of A, although the possibility exists, is attributed to its aromatic nature.

The observations of Gastaldi and Princivalle (91, 178) may also be explained on the basis of these hypotheses. When 2,5-diphenyl(or dimethyl)-6-hydroxypyrazine was treated in a sealed tube for 10 hr. at 100°C. with potassium hydroxide and methyl iodide in methanol, 1,4-dimethyl-2,5-diphenyl(or dimethyl)-6-keto-1,6-dihydropyrazinium iodide was obtained.



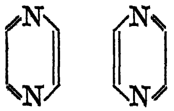
R = phenyl or methyl.

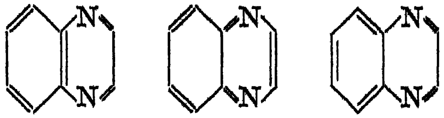
The 1,2,4,5-tetramethyl-6-keto-1,6-dihydropyrazinium iodide, when treated with alkali, may further react to form the anhydro base 1,2,4-trimethyl-5-methylene-6-keto-1,6,4,5-tetrahydropyrazine, analogous to the behavior of 1,2,2,3,5,5,6-heptamethyldihydropyrazinium iodide (C). It was also noted that 1,2,5-trimethylpyrazinium iodide did not form an anhydro base under similar circumstances but underwent another change. This we have seen was identified by Aston as the formation of the unstable dimeric ether.

Before closing our discussion of the quaternary salts we should mention a reaction of these compounds which has received much attention in the pyridine series, *viz.*, the migration of *N*-substituted radicals. An example of this type has been recorded by Brandes and Stoehr (24). They observed that when the methobromide of dimethylpyrazine was heated in a sealed tube, trimethylpyrazine and tetramethylpyrazine were obtained. The formation of the latter compound indicates that the migration is intermolecular, in accord with our knowledge of similar migrations in the pyridine series.

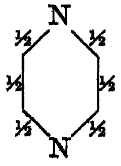
Bergstrom and Ogg (17) have made an interesting interpretation of the pyrazine molecule, based on the ammonia system of compounds. Since it contains the linkage —N=CH—CH=N—, pyrazine is formally an ammono glyoxal; the same is of course true of quinoxaline. These concepts were tested by treating these compounds with typical aldehyde reagents in liquid ammonia. Quinoxaline formed aldehyde addition compounds with bisulfite, hydrogen cyanide, and the Grignard reagent, whereas pyrazine, if it reacted at all, yielded undefined products.

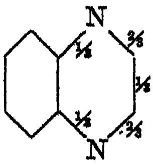
An explanation of this behavior may be formulated on the basis of resonance phenomena similar to Pauling's treatment of benzene and the polynuclear hydrocarbons (171). Considering the following resonating structures involving double bond shifts:


Pyrazine


Quinoxaline

we can assign values for double-bond character:

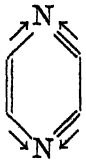




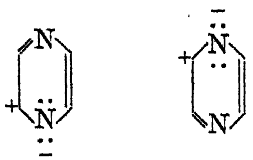
In the case of quinoxaline, therefore, the —N—C— bonds in the 1,2- and 3,4-positions possess more double-bond character than those in the other positions of either quinoxaline or pyrazine. Hence, we should expect these positions to exhibit greater aldehydic properties, a conclusion which is in conformity with the observed facts.

C. DIRECT NUCLEAR SUBSTITUTION

The structure of the pyrazine molecule seems particularly unsuited for direct substitution with the usual electrophilic reagents. In addition to the inductive effects of the negative nitrogen atoms which reduce the electron density of all carbons equally.



the deactivation is intensified by the effects of such resonating structures as:



Of the three diazines, it is only pyrazine which is rendered so thoroughly inactive, for in pyrimidine and pyridazine there is at least one carbon atom which is not adjacent to both nitrogens.

However, just as the nucleus is deactivated with respect to attempted electrophilic substitutions, so may substitutents on the ring be activated for certain reactions. Thus, we may point to the relatively large values for the dissociation constants of two of the carboxylic acids as determined by Sausville and Spoerri (186) from the half-neutralization points of electrometric titrations with the hydrogen electrode: pyrazinoic acid, $K_a = 1.2 \times 10^{-3}$; 2,3-pyrazinedicarboxylic acid, $K_1 = 1.7 \times 10^{-3}$, K_2 "too weak to measure." Further evidence will be presented in the discussions of the ammonolysis, hydrolysis, and alcoholysis of the nuclear-substituted halides and the condensation of alkyl derivatives with aldehydes.

1. *Halogenation*

Until recent times all attempts at halogenation, as shown in table 15, were unsatisfactory, giving rise to products of varying stability and uncertainty of structure.

The first successful direct halogenations of pyrazine have been reported in patents issued within the past year to Sayward (187) and Winnek (231). Both employed vapor-phase reactions reminiscent of the methods used for the halopyridines. The former simply passed a mixture of pyrazine and water vapors with chlorine through a tube at 400°C.; whereas the latter diluted the pyrazine vapor with nitrogen instead of water, and was able to operate at lower temperatures by using a copper–cuprous chloride catalyst. A patent just issued to Dixon and Sayward (64a) claims the preparation of bromopyrazine from the reaction between pyrazine hydrochloride (or hydrobromide) and bromine in carbon tetrachloride at 150–250°C. for 3½ hr.

Certain substituents on the ring apparently modify the character sufficiently to permit a successful direct halogenation under ordinary conditions. Thus Ellingson and Henry (69) found that whereas they could obtain nothing but tars from the action of bromine on aminopyrazine, the same reagent was successful in converting 2-amino-3-carbomethoxypyrazine to 2-amino-3-carbomethoxy-5-bromopyrazine in 90 per cent yields.

Almost simultaneously with the introduction of these direct methods, Erickson and Spoerri (76, 77) prepared pure chloropyrazine by the treatment of hydroxypyrazine with a mixture of phosphorus pentachloride and phosphorus oxychloride. When the corresponding phosphorus bromides were employed, one of the dibromopyrazines was obtained in addition to the monobromo derivative. More recently, McDonald and Ellingson (149) treated 2-amino-3-hydroxypyrazine with phosphorus oxychloride under pressure to obtain 2-amino-3-chloropyrazine.

TABLE 15

Early attempts at the direct halogenation of pyrazines

PYRAZINE	REAGENT	PRODUCT	REFERENCES
Pyrazine	Bromine	Unstable addition compound	(209)
Pyrazine	Iodine and potassium iodide	"Periodide"; structure undetermined	(209)
Trimethylpyrazine	Iodine and potassium iodide	"Periodide"; structure undetermined	(24)
Tetramethylpyrazine	Bromine	Unstable addition compound	(99)
2,5-Dimethylpyrazine	Bromine	Unstable addition compound	(203, 205)
2,5-Dimethylpyrazine	Iodine and potassium iodide	"Periodide"; structure undetermined	(203)
3-Keto-2,5-diphenyl-3,4-dihydropyrazine	Bromine in glacial acetic acid	Fairly stable crystalline monobromo derivative; structure not proven	(109, 153)

Durio and Bissi (66) found that when they subjected 1,4-dioxotetraphenylpyrazine to the action of phosphorus pentachloride, deoxygenation occurred together with the halogenation of one of the phenyl groups to yield 2-chlorophenyl-3,5,6-triphenylpyrazine.

2. *Nitration*

Minovici (109, 153) apparently obtained a nitro derivative of 3-keto-2,5-diphenyl-3,4-dihydropyrazine by the action of fuming nitric acid. The structure of this compound was not proven. Generally when aryl-substituted pyrazines are treated with nitric acid, it is found, as expected, that if nitration does take place, it is the aryl nucleus which is primarily affected. Thus 2,5-di[*m*-nitrophenyl]pyrazine was obtained from the nitration of diphenylpyrazine (163). Similar treatment of tetraphenylpyrazine resulted in the formation of tetra-(nitrophenyl)pyrazine (30), which was capable of being reduced to tetra(aminophenyl)pyrazine by tin and hydrochloric acid (190).

3. *Sulfonation*

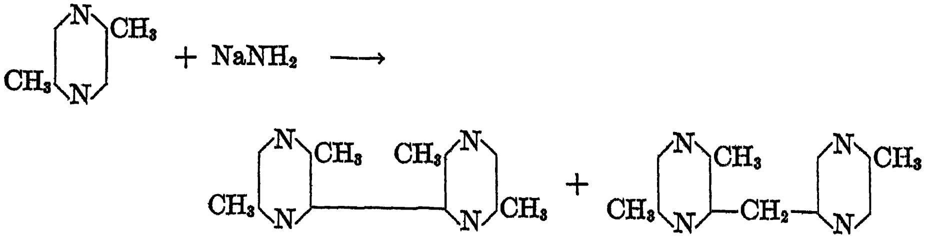
The subject of sulfonation may be dismissed with the brief statement that no pyrazinesulfonic acids have been reported. Some idea of the resistance of the pyrazine nucleus to sulfuric acid may be gained from the fact that lumazine may be cleaved to aminopyrazine by heating in 100 per cent sulfuric acid at 250°C. There is little apparent detriment to the aminopyrazine, since it may be recovered in 80 per cent yields (227).

4. *Amination*

Since those electronic conditions which discourage electrophilic substitutions aid the attack of nucleophilic reagents, one would anticipate favorable results from a direct amination with sodium amide. When this reaction is carried out in the pyridine series, it is observed that it is the 2- and 4-positions which are first substituted, contrary to the action of the electrophilic reagents which preferentially attack the electron-richer 3-position.

Bergstrom and Ogg (17) attempted to react pyrazine with potassium amide in liquid ammonia, but they could isolate no products. However, a patent has just been obtained by Crossley and English (51), in which they claimed to have modified the conditions sufficiently to permit successful results.

Chichibabin and Schukina (41) treated sodium amide with 2,5-dimethylpyrazine in xylene, and obtained only 10 per cent yields of 3-amino-2,5-dimethylpyrazine. In addition, coupling induced by the aminating agent occurred, as in the pyridine series (168), yielding bipyrazyls:



With dimethylaniline as the solvent, Joiner and Spoerri (116, 117) were able to raise the yield of 3-amino-2,5-dimethylpyrazine to 35 per cent.

by the Hofmann degradation of the acid amides. This will be further considered in Section D,4,d, dealing with the properties of the acid amides.

5. *Cyanation*

Another example of an anionoid reagent successful in the pyrazine series is potassium cyanide. Mason *et al.* (147) refluxed 2,3-diphenyl-5,6-dihydropyrazine with potassium cyanide in 80 per cent ethanol. After preliminary oxidation to the pyrazine, cyanation followed by partial hydrolysis occurred, resulting in the amide of 2,3-diphenyl-5-pyrazinoic acid. 2,3-Dimethoxyphenyl-5,6-dihydropyrazine yielded the corresponding amide under similar treatment.

D. REACTIONS OF RING SUBSTITUENTS

1. *Hydrolysis, ammonolysis, and alcoholysis of halopyrazines*

As previously mentioned, the halopyrazines are fairly reactive. Bromopyrazine gives a positive halogen test with alcoholic silver nitrate (76). This resembles the behavior of the 2- and 4-halopyridines. Spoerri and Erickson (76, 77) hydrolyzed chloropyrazine to hydroxypyrazine by heating with aqueous alkali in a sealed tube for 7 hr. Similarly, the ammonolysis of chloropyrazine was accomplished with concentrated ammonia, to produce 2-aminopyrazine (76, 77, 232). 2,3-Diaminopyrazine was obtained by the analogous treatment of 2-amino-3-chloropyrazine with ammonia (149).

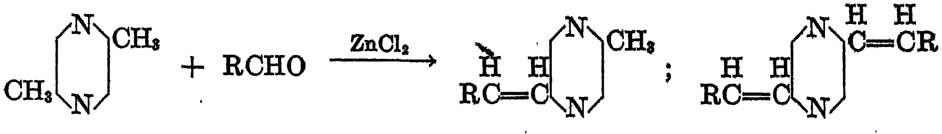
Metanilamidopyrazine and orthanilamidopyrazine were both prepared by treating chloropyrazine with either metanilamide or orthanilamide, in nitrobenzene in the presence of potassium carbonate (72). Winnek (232a) has patented the preparation of 2-N^4-[acetylsulfanilamido]pyrazine from chloropyrazine and N^4-acetylsulfanilamide by a similar method.

An example of the alcoholysis of a halopyrazine is the preparation of ethoxypyrazine by refluxing chloropyrazine in an alcoholic solution of sodium ethoxide (76, 77).

2. *Condensation of alkylpyrazines with aldehydes*

The enhanced activity of the alkyl groups on the pyrazine nucleus may be illustrated by the condensation of the methyl derivatives with various aldehydes. This reaction is entirely analogous to that of other activated methyl groups such as occur in 2,4-dinitrotoluene, the α- and β-picolines, and the 2- and 4-alkylquinolines.

Franke (85) published the results of an investigation in which he condensed various aldehydes with 2,5-dimethylpyrazine in the presence of zinc chloride. Depending upon the conditions and the particular aldehyde, one and/or both methyl groups were found to react:



A summary will be found in table 16.

Riesz (182) extended this work to the quaternary salts of the methylpyrazines, employing a piperidine catalyst. The condensation products were stated to have the property of desensitizing silver halide emulsions. The compounds specifically prepared in the patent also appear in table 16.

3. *Hydrolysis of aminopyrazines*

An example of this reaction which also depends on activation by the ring nitrogens has been furnished by Weijlard, Tishler, and Erickson (227). By heating 2-amino-3-pyrazinoic acid with 20 per cent sodium hydroxide in a sealed tube they were able to obtain 2-hydroxy-3-pyrazinoic acid.

TABLE 16

Condensation of methylpyrazines with aldehydes

REACTANTS	PRODUCTS	REFERENCES
2,5-Dimethylpyrazine and benzaldehyde	2-Methyl-5-styrylpyrazine and 2,5-distyrylpyrazine	(85)
2,5-Dimethylpyrazine and chloral	2,5-Bistrichloropropenylpyrazine	(85)
2,5-Dimethylpyrazine and *p*-toluylaldehyde	2,5-Di(*p*-methylstyryl)pyrazine	(85)
2,5-Dimethylpyrazine and anisaldehyde	2,5-Di(*p*-methoxystyryl)-5-methylpyrazine	(85)
2,5-Dimethylpyrazine and *p*-nitrobenzaldehyde	2-*p*-Nitrostyryl-5-methylpyrazine	(85)
1,2,5-Trimethylpyrazinium nitrate and benzaldehyde	1,5-Dimethyl-2-styrylpyrazinium nitrate	(182)
1-Ethyl-2,5-dimethylpyrazinium iodide and *p*-dimethylaminobenzaldehyde	1-Ethyl-5-methyl-2(4′-dimethylaminostyryl)pyrazinium iodide	(182)
1,2,5-Trimethylpyrazinium iodide and *p*-methoxybenzaldehyde	1,5-Dimethyl-2(4′-methoxystyryl)-pyrazinium iodide	(182)
1-Ethyl-2,5-dimethylpyrazinium iodide and *m*-nitrobenzaldehyde	1-Ethyl-5-methyl-2(3′-nitrostyryl)-pyrazinium iodide	(182)

4. *Oxidation of side chains—the carboxylic acids and their derivatives*

It thus appears that the standard methods for the introduction of substituents into the benzene nucleus which have been used with some measure of success in such heterocycles as pyridine, quinoline, and even the two other diazines are not at all applicable to the pyrazines under similar circumstances. As was indicated, it is only within the past few years that efforts have been made to modify the techniques suitably to make them practical for use in the pyrazine series.

The classical syntheses of pyrazine derivatives generally hinged on the preparation of the carboxylic acids. These could be treated by usual methods to yield esters, amides, nitriles, etc., which were then made the starting materials for further derivatives, such as the amino, hydroxy, and halogen compounds.

TABLE 17

Oxidation of pyrazine side chains

PYRAZINE DERIVATIVE	OXIDATION PRODUCT	REFERENCES
Methylpyrazine	Pyrazinoic acid	(211)
2,5-Dimethylpyrazine	2,5-Pyrazinedicarboxylic acid and/or 2-methyl-5-pyrazinoic acid, depending on conditions	(75, 93, 129, 206, 208, 210, 214, 238)
2,5-Dimethyl-3-ethylpyrazine	Pyrazinetricarboxylic acid and/or 2,5-dimethyl-3-pyrazinoic acid, depending on conditions	(196, 214)
2,5-Dimethyl-3,6-diethylpyrazine	2,5-Dimethyl-3,6-pyrazinedicarboxylic acid	(119)
Tetramethylpyrazine	Pyrazinetetracarboxylic acid. As the permanganate concentration is reduced there is obtained an increasing quantity of 3,6-dimethyl-2,5-pyrazinedicarboxylic acid	(228, 243)
Phenazine and substituted phenazines	Pyrazinetetracarboxylic acid	(181)
2-Methyl-6-pyrazinoic acid	2,6-Pyrazinedicarboxylic acid	(196)
Copper 2-methylol-5-pyrazinoate	2,5-Pyrazinedicarboxylic acid	(215)
Fructosazine	2,5-Pyrazinedicarboxylic acid	(215)
2,2′,5,5′-Tetramethyl-3,3′-bipyrazyl	2,5-Dimethyl-3-pyrazinoic acid and 2,5-dimethyl-2′-methyl-5′-carboxyl-3,3′-bipyrazyl	(41)
(2,5-Dimethyl-5′-methyl)bipyrazyl-3,3′-methane	2,5-Pyrazinedicarboxylic acid and 2,5-dimethyl-3-pyrazineglyoxalic acid	(41)

We first concern ourselves, therefore, with the carboxylic acids. In Section IV, we have noted a few cases where these compounds were obtained directly as the result of a pyrazine synthesis. However, the most important method for the preparation of the pyrazine acids is dependent on the oxidation of the side chains of those alkyl derivatives that are available (see table 17).

The reagent most commonly selected for this task is dilute (2 per cent) alkaline permanganate. This works very well for quinoxaline and its derivatives, resulting in the formation of the corresponding 2,3-dicarboxylic acids in yields of the order of 70 per cent. However, when applied to the oxidation of other side chains, yields are considerably poorer, the total varying from 5 to 10 per cent, because of additional destruction of the nucleus. Rather futile efforts have been made to improve the situation by the control of variables such as the permanganate and alkylpyrazine concentrations, pH, and reaction time. The use of other oxidants, such as nitric acid and acid dichromate mixtures, has also been of little avail (93, 129).

TABLE 18

Pyrazinecarboxylic acids obtained by the hydrolysis of amides or nitriles

AMIDE OR NITRILE	ACID	CONDITIONS	REFERENCES
2,5-Dimethyl-3,6-pyrazinediamide	2,5-Dimethyl-3,6-pyrazinedicarboxylic acid	Nitrous acid	(91)
2,3-Dimethoxyphenyl-5-pyrazinamide	2,3-Dimethoxyphenyl-5-pyrazinoic acid	Heat with aqueous alkali or acid	(147)
2,3-Diphenyl-5-pyrazinamide	2,3-Diphenyl-5-pyrazinoic acid	Heat with aqueous alkali or acid	(147)
Cyanopyrazine	Pyrazinoic acid	Heat with aqueous alkali	(21)
2,5-Diphenyl-3,6-dicyanopyrazine	2,5-Diphenyl-6-hydroxy-3-pyrazinoic acid	Heat with aqueous alkali	(90)
2,5-Diphenylcyanopyrazine	2,5-Diphenylpyrazinoic acid	Heat with aqueous alkali	(90)

Several of the carboxylic acids have also been obtained by the hydrolysis of the amides or nitriles. These are listed in table 18.

A characteristic property of the pyrazinecarboxylic acids is their reaction with ferrous salts to give reddish-colored compounds. This is also true of the pyridine- and quinolinecarboxylic acids (Skraup's test). It has been established that the necessary condition for a positive test is the group, —C=N, in which the dotted portion signifies the remainder of a cyclic system (139).

a. Decarboxylation

In common with the carboxylic acids of other heterocycles, e.g., the pyridines, in good yields. This has proven a lucrative device for the preparation of derivatives otherwise unobtainable.

There are a number of ways in which the decarboxylation may be effected,

TABLE 19

Decarboxylations of pyrazinecarboxylic acids

PYRAZINECARBOXYLIC ACID	PRODUCT	CONDITIONS	REFERENCES
Pyrazinoic acid	Pyrazine	Sublimation; dibutyl phthalate	(180, 206)
2,3-Pyrazinedicarboxylic acid	Pyrazinoic acid	Vacuum sublimation	(16, 21, 60, 89, 100, 101, 116, 180)
2,3-Pyrazinedicarboxylic acid	Pyrazine	Vacuum sublimation; dibutyl phthalate	(16, 89, 93, 116)
2,5-Pyrazinedicarboxylic acid	Pyrazinoic acid	Sublimation; glacial acetic acid	(238, 243)
2,5-Pyrazinedicarboxylic acid	Pyrazine	Sublimation; glacial acetic acid	(166, 206, 208, 210, 214, 238)
Pyrazinetricarboxylic acid	2,5-Pyrazinedicarboxylic acid; also some 2,6-pyrazinedicarboxylic acid	Boiled in water for 2 to 3 days; dibutyl phthalate	(180, 206, 214)
Pyrazinetetracarboxylic acid	2,5-Pyrazinedicarboxylic acid; also some pyrazine	Fuse dipotassium salt; dibutyl phthalate	(180, 238, 243)
2-Methylpyrazinoic acid	Methylpyrazine	Sublimation	(206, 208, 210)
2,5-Dimethylpyrazine-3,6-dicarboxylic acid	2,5-Dimethylpyrazine	Glacial acetic acid	(206, 208, 210, 214)
5,6-Dimethylpyrazine-2,3-dicarboxylic acid	5,6-Dimethylpyrazine	Glacial acetic acid	(89)
2,5-Dimethyl-6-hydroxy-3-pyrazinecarboxylic acid	2,5-Dimethyl-6-hydroxypyrazine	Fuse sodium salt	(90)
2-Hydroxy-3-pyrazinoic acid	2-Hydroxypyrazine	Carbitol acetate	(227)
2-Amino-5,6-dimethyl-3-pyrazinoic acid	2-Amino-5,6-dimethylpyrazine	80 per cent sulfuric acid	(227)
2-Amino-5,6-diphenylpyrazinoic acid	2-Amino-5,6-diphenylpyrazine	80 per cent sulfuric acid	(227)

TABLE 19—*Continued*

PYRAZINECARBOXYLIC ACID	PRODUCT	CONDITIONS	REFERENCES
2-Amino-6-methyl-3-pyrazinoic acid	2-Amino-6-methylpyrazine	80 per cent sulfuric acid	(227)
2,5-Diphenyl-6-hydroxy-3-pyrazinoic acid	2,5-Diphenyl-6-hydroxypyrazine	Fuse sodium salt	(90)
2,5-Diphenylpyrazinoic acid	2,5-Diphenylpyrazine	Fuse sodium salt	(90)
2-Amino-3-pyrazinoic acid	2-Aminopyrazine	Fuse	(89, 227)
5-Methylpyrazine-2,3-dicarboxylic acid	6-Methyl-2-pyrazinoic acid	Vacuum sublimation	(135)
2,3-Diphenyl-5-carboxypyrazine-6-*o*-benzoic acid	2,3,5-Triphenylpyrazine	Distill over calcium oxide	(48)
Pyrazine-3-phenyl-5-carboxy-6-*o*-benzoic acid	2,5-Diphenylpyrazine	Distill over calcium oxide	(46)

as is indicated in table 19. The general procedure is to reflux the acid in glacial acetic acid. This has been varied by employing other media such as carbitol acetate, 80 per cent sulfuric acid, and dibutyl phthalate—the major effect being the increase in temperature due to the higher boiling point. Some acids undergo this reaction so easily that it may be conducted in boiling water. Sublimation has also been found to be effective, and this process was further improved for some cases by operating under reduced pressures. Occasionally, heating the alkali salt to the state of fusion was sufficient to induce decarboxylation.

An example of the decarboxylation of a cyano derivative was furnished by Gastaldi (90), who found that phosphorus and hydrogen iodide converted 2,5-diphenyl-3,6-dicyanopyrazine into 2,5-diphenyl-3,6-dihydropyrazine.

b. Esters

Esterification with methanol and ethanol takes place smoothly when carried out in the customary fashion. Generally, the acid is dissolved in an excess of the alcohol which contains either hydrogen chloride or sulfuric acid as a catalyst. After warming and/or allowing to stand for a period of time, the esters usually crystallize out when the solvent is partially evaporated. The yields are good, ranging from 80–90 per cent.

These derivatives have been largely utilized for the preparation of the acid amides.

The β-diethylaminoethyl ester of 2-amino-3-pyrazinoic acid was prepared by the reaction of silver 2-amino-3-pyrazinoate with diethylaminoethyl chloride (73). It was thought that this compound might possess anesthetic activity because of its resemblance to procaine.

TABLE 20

Esterifications of pyrazinecarboxylic acids

PYRAZINECARBOXYLIC ACID	PRODUCT	CONDITIONS	REFERENCES
2,3-Pyrazinedicarboxylic acid	Dimethyl 2,3-pyrazinedicarboxylate	Methanol; hydrogen chloride	(73, 89, 166)
5-Methylpyrazine-2-pyrazinoic acid	5-Methyl-2-methyl pyrazinoate	Methanol; hydrogen chloride	(227)
Pyrazinoic acid	Methyl pyrazinoate	Methanol; hydrogen chloride	(21, 144)
Pyrazinoic acid	Ethyl pyrazinoate	Ethanol; hydrogen chloride	(116, 117)
2,5-Pyrazinedicarboxylic acid	Dimethyl 2,5-pyrazinedicarboxylate	Methanol; hydrogen chloride	(75, 129)
2-Amino-3-pyrazinoic acid	2-Amino-3-methyl pyrazinoate	Methanol; sulfuric acid	(70)
5,6-Dimethyl-2,3-pyrazinedicarboxylic acid	Dimethyl 5,6-dimethyl-2,3-pyrazinedicarboxylate	Methanol; hydrogen chloride	(89)
Pyrazinetetracarboxylic acid	Tetraethyl pyrazinetetracarboxylate	Ethanol; hydrogen chloride	(37)
3-Hydroxy-2-pyrazinoic acid	3-Hydroxy-2-methyl pyrazinoate	Methanol; sulfuric acid	(139)

c. Acid chlorides

The literature records the preparation of two pyrazine acid chlorides. These are pyrazinoyl chloride (57) and the diacid chloride of 2,5-pyrazinedicarboxylic acid (75), both of which were made by the action of phosphorus pentachloride on the free acids. Thionyl chloride was found to be an ineffective reagent for this transformation (75).

d. Amides and hydrazides

The acid chlorides and esters have been used to prepare the amides and hydrazides by the standard procedure of treating them with ammonia, hydrazine, or their derivatives. Yields are good: 80–90 per cent. For some of the higher

N-substituted amides, it was found sufficient to reflux the pyrazine acid directly with the amine in xylene (18).

The hydrolytic action of 95 per cent sulfuric acid on nitriles has also served for the preparation of the amides.

The appearance of the large number of these compounds may be attributed in part to their importance as intermediates for the Hofmann and Curtius degradations. The majority of them, however, have been synthesized by Dalmer and Walter (56) because of their purported physiological behavior as analeptics.

TABLE 21

Amides and hydrazides

PYRAZINE REACTANT	PRODUCT	CONDITIONS	REFERENCES
Methyl or ethyl pyrazinoate	Pyrazinamide	Alcoholic ammonia	(21, 116)
Dimethyl 2,3-pyrazinedicarboxylate	2,3-Pyrazinediamide	Alcoholic ammonia	(73, 89, 166)
Dimethyl 2,5-pyrazinedicarboxylate	2,5-Pyrazinediamide	Alcoholic ammonia	(75, 88, 129)
2-Amino-3-carbomethoxypyrazine	2-Amino-3-pyrazinamide	Alcoholic ammonia	(70)
2-Methyl-5-carbomethoxypyrazine	2-Methyl-5-pyrazinamide	Alcoholic ammonia	(227)
2,3-Pyrazinedicarboxylic acid	*N*,*N'*-Dibenzyl-2,3-pyrazinediamide	Benzylamine refluxed with the diacid in xylene	(18)
2,3-Pyrazinedicarboxylic acid	*N*,*N'*-Diamyl-2,3-pyrazinediamide	Amylamine refluxed with the diacid in xylene	(18)
Diethyl 2,5-pyrazinedicarboxylate	2,5-Pyrazine dihydrazide	Hydrazine hydrate	(75)
3-Hydroxy-2-carbomethoxypyrazine	3-Hydroxy-2-pyrazine amide	Alcoholic ammonia	(149)
2,5-Dimethyl-3,6-dicyanopyrazine	2,5-Dimethyl-3,6-pyrazinediamide	Heat with sulfuric acid	(91)
2-Amino-3-cyanopyrazine	2-Amino-3-pyrazinamide	Heat with sulfuric acid	(70)

The following derivatives were claimed in the patent issued to Dalmer and Walter (56). They were synthesized by the action of ethyl pyrazinoate or pyrazinoyl chloride on ammonia, hydrazine, or their derivatives: *N*-β-amino-

ethylpyrazinamide; *N*-β-hydroxyethylpyrazinamide, *N*-butylpyrazinamide, *N*,*N*-dibenzylpyrazinamide, *N*,*N*-di(β-hydroxyethyl)pyrazinamide, *N*,*N*-dibutylpyrazinamide, *N*,*N*-diisopropylpyrazinamide, *N*,*N*-dimethylpyrazinamide, *N*,*N*-dipropylpyrazinamide, *N*-ethylpyrazinamide, *N*-heptylpyrazinamide, *N*-methylpyrazinamide, *N*-methylbenzoylpyrazinamide, *N*-methylpyrazinhydrazide, *N*-phenylpyrazinhydrazide, pyrazinanilide, pyrazinhydrazide.

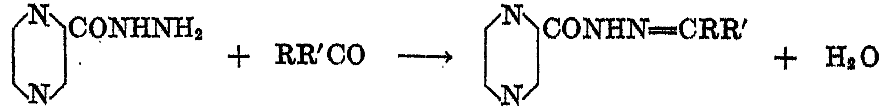
Several sulfanilamide derivatives have been prepared by reacting pyrazinoyl chloride with the appropriate sulfanilamide (57). These are illustrated in table 22.

TABLE 22

Pyrazinoyl sulfanilamide derivatives (57)

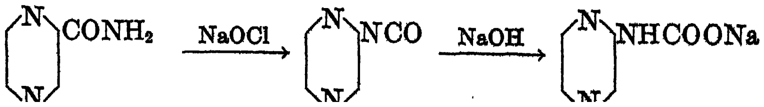
REACTANTS	PRODUCTS
Pyrazinoyl chloride and sulfanilamide	N^4-Pyrazinoylsulfanilamide
N^4-Pyrazinoylsulfanilamide and acetic anhydride	N^1-Acetyl-N^4-pyrazinoylsulfanilamide
N^4-Pyrazinoylsulfanilamide and pyrazinoyl chloride	N^1,N^4-Dipyrazinoylsulfanilamide
N^4-Acetylsulfanilamide and pyrazinoyl chloride	N^4-Acetyl-N^1-pyrazinoylsulfanilamide
N^4-Acetyl-N^1-pyrazinoylsulfanilamide and sodium hydroxide	N^1-Pyrazinoylsulfanilamide

The hydrazino groups of the hydrazides are reactive, readily combining with aldehydes and ketones to form the alkylidenepyrazinoyl hydrazides (55):



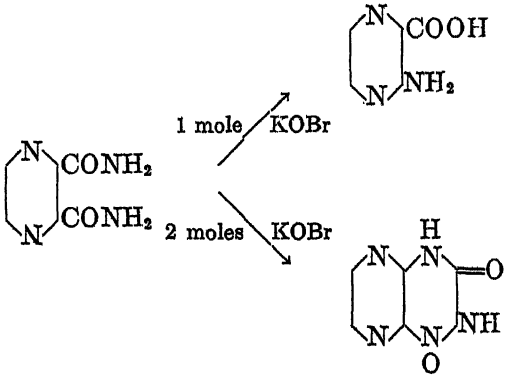
The following derivatives were thus prepared, pyrazinoylhydrazide being condensed with the appropriate carbonyl compound: isopropylidenepyrazinoyl hydrazide (R = R′ = CH_3), ethylidenepyrazinoyl hydrazide (R = CH_3, R′ = H), and trichloroethylidenepyrazinoyl hydrazide (R = CCl_3, R′ = H).

With the few exceptions already noted, all the reported aminopyrazines have been prepared by the Hofmann degradation. There are several points of interest with regard to this reaction. For example, the pyrazine nucleus may exert a stabilizing influence on some of the intermediates. Thus, the sodium salt of the intermediate carbamate could be isolated when the reaction was carried out with pyrazinamide (101):



In addition to the preparation of aminopyrazine (21, 101, 116), the method has been successfully applied to two other pyrazine monoamides, 2-methyl-5-pyrazinamide (227) and 2-hydroxy-3-pyrazinamide (149), for the preparation of the corresponding 2-methyl-5-aminopyrazine and 2-hydroxy-3-aminopyrazine.

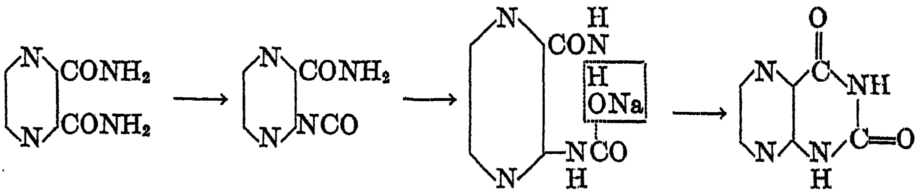
Gabriel and Sonn (89) investigated the diamide of the ortho diacid. They discovered that, although 1 mole of hypobromite reacted with 1 mole of the diamide to give the expected 2-amino-3-pyrazinoic acid, 2 moles of the hypobromite effected cyclization:



A similar course of events was observed in the pyridine series.

The condensed heterocycle was subsequently synthesized in other ways and given the name of lumazine. As has been already noted, use has been made of this compound and its derivatives in the synthesis of other pyrazines.

The formation of lumazine may be accounted for on the basis of the intermediate formation of a carbamate which is sufficiently stable under the conditions of the reaction to undergo intramolecular condensation:



It would seem, therefore, that the yields of lumazine and 2-amino-3-pyrazinoic acid are functions of the relative rates at which the carbamate group reacts either intramolecularly with the amide group or intermolecularly with the alkaline medium. Furthermore such rates, from the observations of Gabriel and Sonn, apparently depend upon the concentration of the hypobromite.

Ellingson, Henry, and McDonald (70) were unsuccessful in an attempt to prepare the 2,3-diamine by the Hofmann degradation of 2-amino-3-pyrazinecarboxamide. When the acetyl derivative was employed, two unexpected products were obtained which were not identified.

Further evidence of the stabilizing effect of the pyrazine nucleus is found in the work of Erickson and Spoerri (75) on 2,5-pyrazinedicarboxamide. After failing to obtain the diamine by the normal Hofmann degradation, they prepared the intermediate compounds, using the Curtius reactions. In this way they could determine at which stage the sequence was halted. The diazide of 2,5-pyrazinedicarboxylic acid was synthesized by treating the dihydrazide with nitrous acid (it could not be obtained by reacting the diacid chloride with sodium azide). After either heating the diazide in benzene or treating it with ethanol, the diisocyanate or the diurethan was respectively obtained.

Now, in the normal Hofmann and Curtius reactions the isocyanate which is formed immediately hydrolyzes under the conditions to the amine. In this case, however, the diisocyanate (and the diurethan) was found to be extremely stable toward both alkaline and acid hydrolysis. Heating the diisocyanate (or diurethan) with fuming hydrochloric acid at 210°C. in a sealed tube or fusing it with solid potassium hydroxide had slight effect. It is of interest to add that the diamine sought could not be prepared according to Gabriel's method, for the diphthalimide which was made by the action of phthalic anhydride on the diazide proved incapable of hydrolysis.

In addition to the Hofmann degradation, the amides undergo the other reactions characteristic of these compounds. An example of the Bouveault reaction with nitrous acid is given in table 18.

The dehydration of the amides to the nitriles requires the most vigorous dehydrating agents. The syntheses of cyanopyrazine (21, 60) and 3-amino-2-cyanopyrazine (70) were achieved only by distilling the corresponding amides with phosphorus pentoxide.

Gabriel and Sonn (89, 166) formed the imide of 2,3-pyrazinedicarboxylic acid by heating the 2,3-diamide in vacuum.

The 2,5-dicarboxamide was observed to be quite refractory toward the standard dehydrating agents (129). After refluxing the compound in acetic anhydride in an attempt to dehydrate one of the amide groups, it was found that acetylation had taken place after 48 hr., forming the diacetyldiamide of 2,5-pyrazinedicarboxylic acid. To obtain the 2,5-dicyanopyrazine it was necessary to reflux the diamide with phosphorus pentoxide in nitrobenzene, the solvent being necessary to avoid extensive decomposition.

e. Cyanopyrazines

There is little that need be said concerning the cyanopyrazines. They have been prepared by the dehydration of amides, by direct cyanation, and by the cyclization of certain nitriles. The hydrolytic reactions have already been discussed under the carboxylic acids and amides. An investigation was made to determine whether these compounds undergo any of the typical addition reactions (21, 60). Cyanopyrazine was successfully converted into the pyrazinoyl methyliminoether hydrochloride by the action of hydrogen chloride in methanol.

This addition compound was then treated with ammonia to form the pyrazinoylamidine hydrochloride.

f. Acid anhydrides—pyrazine "phthaleins"

The ortho diacid in many ways resembles phthalic acid. Thus it is easily converted to the anhydride by heating with either acetic anhydride or thionyl chloride.

De and Dutta (59) have synthesized a number of the phthalein-type dyes by condensing the *o*-pyrazinedicarboxylic acid with various phenols and amines. The use of such condensing agents as zinc chloride was not essential, but in some cases did improve the yields. Compared with the corresponding benzene derivatives, the pyrazine "phthaleins" were found to possess greater intensity of color but less fluorescence.

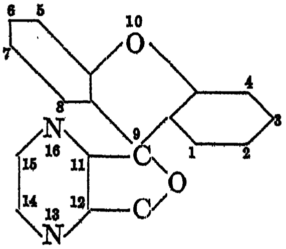
Table 23 summarizes the compounds which have been thus prepared. In order to facilitate the naming of these compounds we shall use the following skeleton, which will be termed pyrazine dicarboxylein in accordance with the usage of De and Dutta:

TABLE 23

Pyrazine dicarboxyleins (59)

COMPOUND CONDENSED WITH 2,3-PYRAZINEDICARBOXYLIC ACID	PRODUCT
Resorcinol	3,6-Dihydroxypyrazine dicarboxylein*
Orcinol	2,7-Dimethylpyrazine dicarboxylein
Phloroglucinol	1,3,6,8-Tetrahydroxypyrazine dicarboxylein
m-Aminophenol	3,6-Diaminopyrazine dicarboxylein
N-Dimethylamino-*m*-phenol	3,6-Di(*N*-dimethylamino)pyrazine dicarboxylein
N-Diethylamino-*m*-phenol	3,6-Di(*N*-diethylamino)pyrazine dicarboxylein
m-Phenylenediamine	3,6-Diamino-10-iminopyrazine dicarboxylein

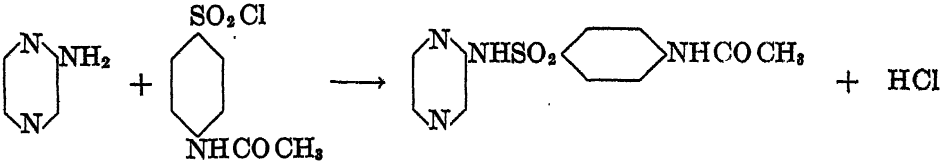
* The bromination of 3,6-dihydroxypyrazine dicarboxylein resulted in the formation of 2,4,5,7-tetrabromo-3,6-dihydroxypyrazine dicarboxylein.



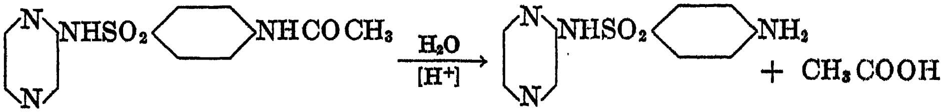
5. Aminopyrazines

We have seen that the aminopyrazines may be prepared directly by amination, by ammonolysis of the halopyrazines, or more widely by the Hofmann degradation of the amides.

The most important practical application of these substances has been in the syntheses of various sulfapyrazines. Generally, the aminopyrazine is first treated with acetylsulfanilyl chloride:

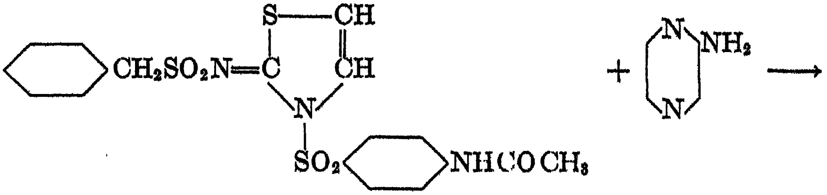


The N^4-acetylsulfanilamidopyrazine is then hydrolyzed to split off the protecting acetyl group:

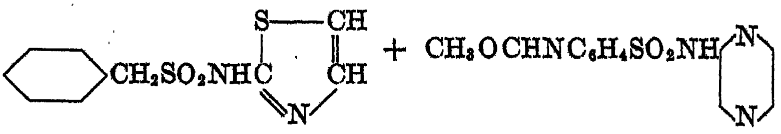


In table 24 there will be found a summary of the sulfanilamidopyrazines that have been made in this manner.

A novel method of preparing sulfanilamidopyrazine from aminopyrazine is that reported in the patent of Hartmann, Cueni, Druey, and von Meyenberg (102a). It is stated that when aminopyrazine is reacted with 2-[benzylsulfonimido]-3-[*p*-acetylaminobenzenesulfonyl]thiazoline (I) an exchange takes place to form acetylsulfanilamidopyrazine (II) and 2-benzylsulfonamidothiazole (III).



I



III II

The latter may be treated with *p*-acetylsulfanilyl chloride to regenerate the sulfanilamide donor (I). Such a technique further suggests itself for the preparation of those sulfanilamidopyrazine derivatives which may be otherwise inaccessible (*vide infra*).

While investigating the sulfanilamidopyrazines, Ellingson and his coworkers (70, 149) discovered certain peculiarities evidently due to the electronic effects of the two ring nitrogens. They found it impossible, for example, to prepare

3-sulfanilamido-2-pyrazinoic acid directly by the condensation of 3-amino-2-pyrazinoic acid with sulfanilyl chloride. To obtain the desired product it was necessary to proceed by way of the hydrolysis of 3-sulfanilamido-2-pyrazinamide (table 24).

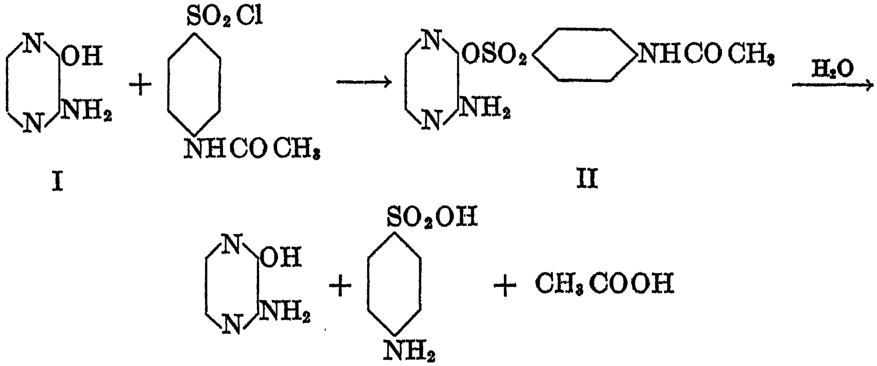
TABLE 24

Sulfanilamidopyrazines

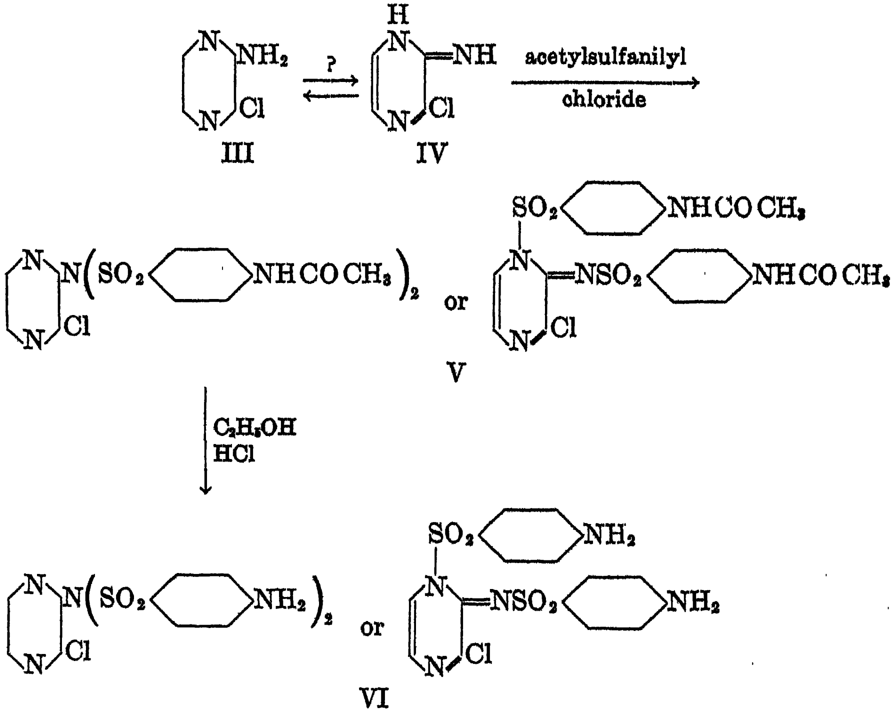
AMINOPYRAZINE	N^4-ACETYLSULFANILAMIDOPYRAZINE	SULFANILAMIDOPYRAZINE	REFERENCES
2,5-Dimethyl-3-amino-aminopyrazine	3-[N^4-Acetylsulfanilamido]-2,5-dimethylpyrazine	3-Sulfanilamido-2,5-dimethylpyrazine	(116)
Aminopyrazine	2-[N^4-Acetylsulfanilamido]pyrazine	2-Sulfanilamidopyrazine	(68, 70, 179, 185, 186)
2-Amino-5,6-dimethylpyrazine	2-[N^4-Acetylsulfanilamido]-5,6-dimethylpyrazine	2-Sulfanilamido-5,6-dimethylpyrazine	(227)
2-Amino-5,6-diphenylpyrazine	2-[N^4-Acetylsulfanilamido]-5,6-diphenylpyrazine	2-Sulfanilamido-5,6-diphenylpyrazine	(227)
2-Amino-6-methylpyrazine	2-[N^4-Acetylsulfanilamido]-6-methylpyrazine	2-Sulfanilamido-6-methylpyrazine	(227)
2-Amino-5-methylpyrazine	2-[N^4-Acetylsulfanilamido]-5-methylpyrazine	2-Sulfanilamido-5-methylpyrazine	(227)
2-Amino-5(or 6)-phenylpyrazine	2-[N^4-Acetylsulfanilamido]-5 (or 6)-phenylpyrazine	2-Sulfanilamido-5(or 6)-phenylpyrazine	(227)
2-Amino-3-carbomethoxypyrazine	2-[N^4-Acetylsulfanilamido]-3-carbomethoxypyrazine	2-Sulfanilamido-3-carbomethoxypyrazine	(90)
3-Amino-2-pyrazinamide	3-[N^4-Acetylsulfanilamido]-2-pyrazinamide	3-Sulfanilamido-2-pyrazinamide	(90)

In attempting to prepare the sulfanilamido derivatives of some bifunctionally substituted pyrazines, further interesting behavior was brought to light (149).

When 2-hydroxy-3-aminopyrazine (I) was treated with acetylsulfanilyl chloride the hydroxyl group was preferentially esterified to yield 3-amino-2-pyrazinyl *p*-acetamidobenzenesulfonate (II). However, when it was desired to cleave the protecting acetyl group, this could not be done without hydrolyzing the sulfonic ester linkage as well:


I II

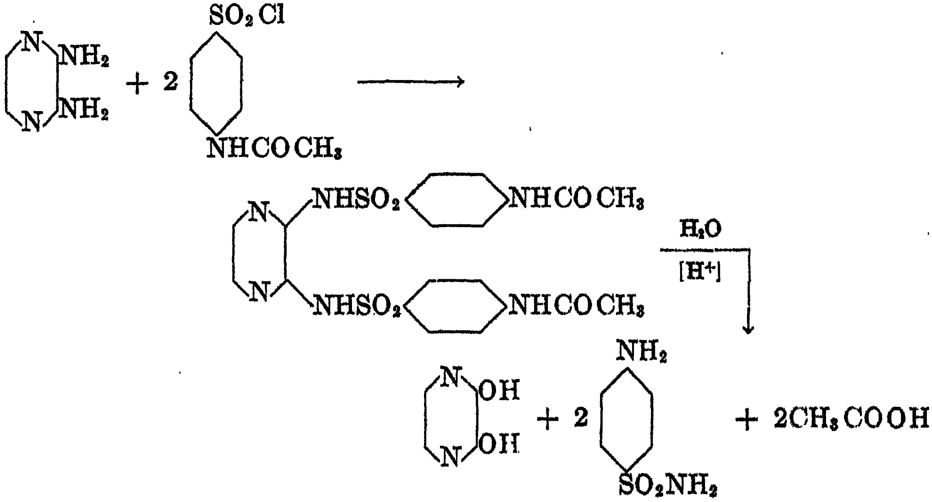
2-Amino-3-chloropyrazine (III) with acetylsulfanilyl chloride yielded a 2-(?)-di-N^4-acetylsulfanilylamido-3-chloropyrazine (V). The question mark appears since there are two possible ways in which the condensation may have occurred: either both acetylsulfanilyl groups reacted with both hydrogens of the amino nitrogen or, with one hydrogen atom on each of two of the nitrogens of the tautomeric imino form (IV).


III IV V VI

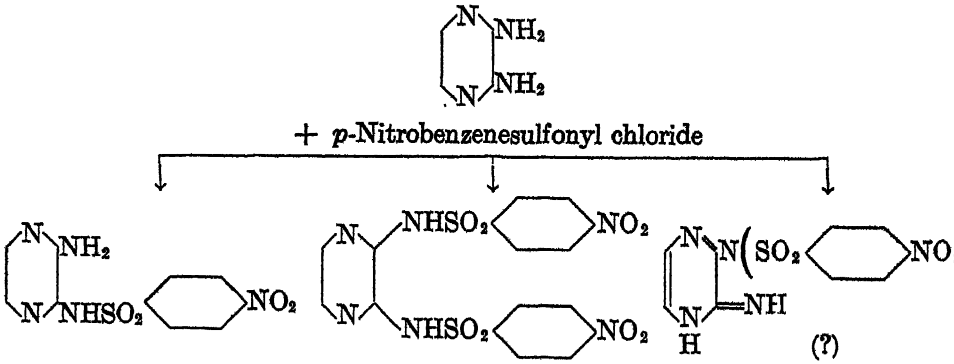
All efforts to induce monocoupling only, by proper adjustment of reagent concentrations, failed. Nor could the monocoupled compound be obtained by the controlled hydrolysis of the dicoupled compound; the only effect was to cleave the protecting acetyl groups, resulting in 2-(?)-disulfanilylamino-3-chloropyrazine (VI).

When acetylsulfanilyl chloride was reacted with 2,3-diaminopyrazine, one hydrogen atom of each of the amino groups was replaced, forming 2,3-di-[N^4-acetylsulfanilamido]pyrazine. The evidence for this was the solubility of the resulting compound in dilute alkali, for the only other alternative, 2-amino-3-di-[N^4-acetylsulfanilamido]pyrazine, would be expected to be more basic because of its free amino group.

Again, attempts to remove the protecting acetyl groups resulted in the cleavage of the sulfanilamido residue.



The action of *p*-nitrobenzenesulfonyl chloride on 2,3-diaminopyrazine resulted in the formation of three substances:



A
2-Amino-3[*p*-nitrobenzenesulfonamido]pyrazine

B
2,3-Di[*p*-nitrobenzenesulfonamido]pyrazine

C
3-Imino-2-di[*p*-nitrobenze sulfonamido]-3,4-dihydr pyrazine

The reduction of B gave several products which could not be characterized.

Because of the weak bascity of the aminopyrazines, the usual methods of diazotization are not very efficient. In order to consummate the reaction it has been found necessary to use nitrosylsulfuric acid in concentrated sulfuric acid. By such means the diazotization of aminopyrazine and 2-amino-3-pyrazinoic acid was accomplished. These diazonium compounds were not isolated, but immediately hydrolyzed into hydroxypyrazine and 2-hydroxy-3-pyrazinoic acid, respectively (76, 77).

6. Hydroxypyrazines

In addition to the above device of hydrolyzing the diazonium compounds, the hydroxypyrazines have been prepared by the hydrolysis of the halopyrazines and the aminopyrazines.

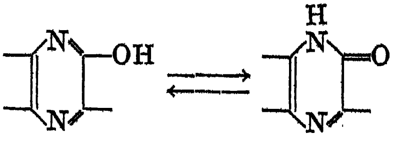
The hydroxypyrazines may be easily esterified by acetic anhydride and benzoyl chloride. These reactions are listed in table 25. For completeness we have also included those pyrazines which contain hydroxyl groups on the side chains.

TABLE 25

Esterification of hydroxypyrazines

REACTANTS	ESTER	REFERENCES
Fructosazine and acetic anhydride	Fructosazine octaacetate	(140)
2-Methylol-5-pyrazinoic acid and acetic anhydride	Acetic ester of 2-methylol-5-pyrazinoic acid	(215)
Hydroxypyrazine and acetic anhydride	Acetoxypyrazine	(76, 77)
Di-*o*-hydroxyphenylpyrazine and benzoyl chloride	Di-*o*-benzoyloxy-2,5-diphenylpyrazine	(224)
o,*o*′,*p*,*p*′-Tetrahydroxy-2,5-diphenylpyrazine	*o*,*o*′,*p*,*p*′-Tetrabenzoyloxy-2,5-diphenylpyrazine	(224)

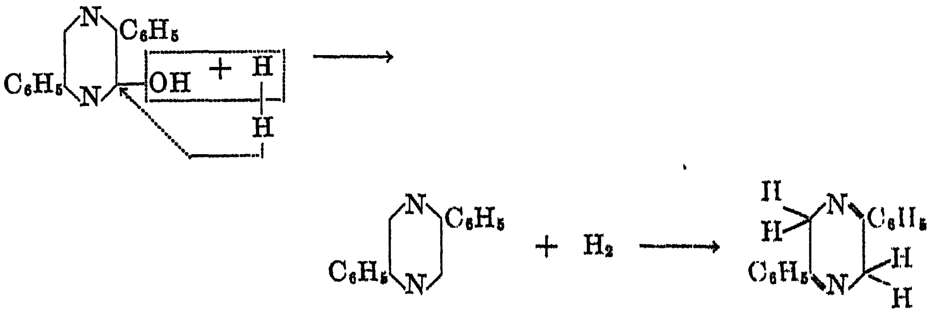
The hydroxypyrazines are formally capable of tautomerizing into ketodihydropyrazines:



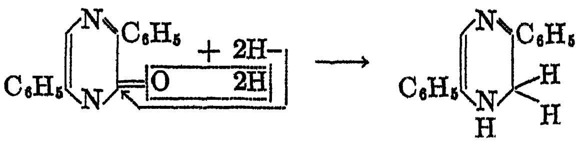
In some instances, for example, 2,5-diphenyl-3-hydroxypyrazine (112), the keto form seems to predominate completely. This conclusion is based on such behavior as the inability of the hydroxyl group to undergo acetylation.

Japp and Knox (112) offered as further evidence the fact that this substance is reduced to 2,5-diphenyl-3,4-dihydropyrazine with phosphorus and hydrogen iodide. Their reasoning was that if the compound existed in the keto modification the oxygen atom would be replaced by the two atoms of hydrogen simultaneously, leading to the 3,4-dihydropyrazine. In the case of the enol however, it was thought that the hydroxyl group would be replaced by hydrogen, forming a pyrazine which could not be further reduced. Now, this is not a strictly valid argument, for it is based upon the premise that phosphorus and hydrogen iodide will not reduce 2,5-diphenylpyrazine. That it could be reduced by this reagent was shown by the work of Gastaldi (90), who obtained 2,5-diphenyl-3,6-dihydropyrazine as the result of such a reaction. The argument of Japp and Knox may be revised, however, to bring it into conformity with the evidence. One must first note that the dihydropyrazines obtained by Japp and Knox and Gastaldi differ in the location of the saturating hydrogen atoms.

It is a plausible assumption that on reduction a hydroxypyrazine will give the same dihydropyrazine as a pyrazine, for it doubtless passes through the latter stage:



On the other hand, a keto pyrazine may yield a different dihydropyrazine, for the carbonyl carbon may be reduced without passing through a pyrazine stage:



Thus, it is seen that the enol form should give a 3,6-dihydropyrazine identical with that obtained by Gastaldi from 2,5-diphenylpyrazine, whereas the keto form should yield a 3,4-dihydropyrazine. Since the latter compound was actually obtained, it must be concluded that the ketonic formulation is correct.

Gastaldi and Princivalle have attacked this problem for the alkylhydroxypyrazines by studying the behavior of these compounds with diazonium salts. It was observed that coupling took place in a normal manner, and this was considered as good evidence that the alkylhydroxypyrazines possessed the enolic configuration.

The reaction was successfully carried out with 2,5-dimethyl-3-hydroxypyrazine and the phenyldiazonium, *o*-tolyldiazonium, and *p*-tolyldiazonium salts to yield the corresponding azo dyes: 3-phenylazo-2,5-dimethyl-6-hydroxypyrazine, 3-*o*-tolylazo-2,5-dimethyl-6-hydroxypyrazine, and 3-*p*-tolylazo-2,5-dimethyl-6-hydroxypyrazine (178). The necessity for the enol form in coupling is further shown by the inability of 1,2,5-trimethyl-6-keto-1,6-dihydropyrazine to couple with diazonium compounds (178).

The coupling of *p*-hydroxypyrazinecarboxylic acids is unusual in that it is accompanied by the elimination of the carboxyl group. Thus 2,5-dimethyl-6-hydroxy-3-pyrazinoic acid, coupled with either benzenediazonium chloride or *p*-tolyldiazonium chloride, yields either 3-phenyl-2,5-dimethyl-6-hydroxypyrazine or 3-*p*-tolylazo-2,5-dimethyl-6-hydroxypyrazine (178).

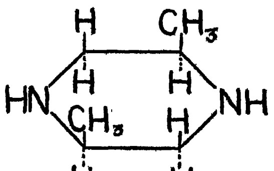
Similar behavior has been observed in the coupling of certain other hydroxycarboxylic acids, among which may be mentioned *α*-hydroxynaphthoic acid, *p*-hydroxybenzoic acid, and acetoacetic ester.

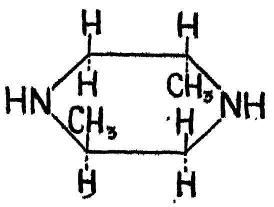
E. REDUCTION OF PYRAZINES

The pyrazine nucleus resembles that of pyridine in its behavior toward reducing agents. They are both much more easily reduced then benzene; in fact, sodium and ethanol, which do not affect benzene, are very effective for the reduction of pyridine and pyrazine to the fully saturated piperidines and piperazines. The analogy with pyridine cannot be carried too far for the other diazines; for, although some pyridazine derivatives may be smoothly reduced, others, including pyridazine itself, are cleaved at the N—N bond.

Other methods have been applied toward the reduction of the pyrazines. Analgams of various types, tin and hydrochloric acid, and catalytic hydrogenation have been employed for the conversion into the piperazines. Red phosphorus and hydrogen iodide is a reagent which will effect partial reduction. This is illustrated in its action on 2,5-diphenylpyrazine, which is reduced to 2,5-diphenyl,3,6-dihydropyrazine (90). Table 26 is a compilation of the reductions carried out in the pyrazine series.

It is perhaps not amiss to point out that in the case of the substituted piperazines the possibilities of geometrical and optical isomerism may arise. Thus for 2,5-dimethylpiperazine *cis* and *trans* forms may result, of which the former would constitute a racemic pair, and the latter an internally compensated meso form.


CIS


TRANS

Kipping (123, 124) has reinvestigated the reduction of tetramethylpyrazine. There are five possible geometric isomers, of which two will be racemic pairs (II to V), and three will be meso forms (I, III, and IV).

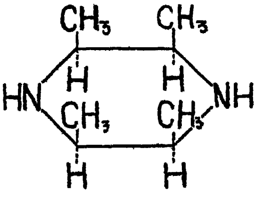
Attempts were made to isolate all five of the isomers by using various reducing agents (table 26) and then separating the products by fractional crystallization.

TABLE 26

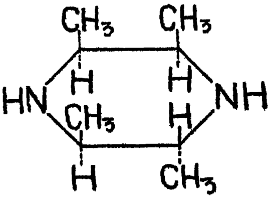
Reduction of pyrazines

PYRAZINE	REDUCED PYRAZINE	REDUCING AGENT	REFERENCES
2,5-Dimethylpyrazine	2,5-Dimethylpiperazine (*cis* and *trans*)	Sodium in ethanol	(177, 207)
2,5-Dimethylpyrazine	2,5-Dimethylpiperazine (*trans*)	Hydrogen over nickel	(94)
Trimethylpyrazine	Trimethylpiperazine (two stereoisomers)	Sodium in ethanol	(207)
Pyrazine	Piperazine	Sodium in ethanol	(238)
2-Methylpyrazine	2-Methylpiperazine	Sodium in ethanol	(211)
2,5-Dimethyl-3-ethylpyrazine	2,5-Dimethyl-3-ethylpiperazine (two stereoisomers)	Sodium in ethanol	(207)
2,3-Diphenylpyrazine	2,3-Diphenylpiperazine (two stereoisomers)	Sodium in amyl alcohol	(146)
2,5-Diphenylpyrazine	2,5-Diphenyl-3,6-dihydropyrazine	Phosphorus and hydrogen iodide	(90)
Tetramethylpyrazine	Tetramethylpiperazine (two stereoisomers)	Sodium in ethanol	(207)
Tetramethylpyrazine	β- and γ-tetramethylpiperazines	Hydrogen over platinum	(123)
Tetramethylpyrazine	α-, β-, and γ-tetramethylpiperazines	Aluminum amalgam in neutral solution	(123)
Tetramethylpyrazine	α- and β-tetramethylpiperazines	Sodium amalgam in neutral solution	(123)
Tetramethylpyrazine	α-, β-, γ-, and δ-tetramethylpiperazines	Tin and hydrochloric acid	(123)
Tetramethylpyrazine methiodide	2,3,4,5,6-Pentamethylpiperazine	Hydrogen over platinum	(124)

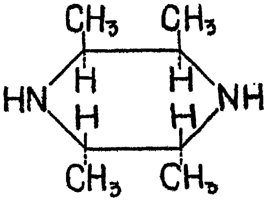
In order to determine the configuration of each of the stereoisomeric fractions, the following approach was used. II and V could be differentiated from the meso forms by the fact that they were capable of resolution. It was further noted that in II the environment of the two ring nitrogen atoms was different,



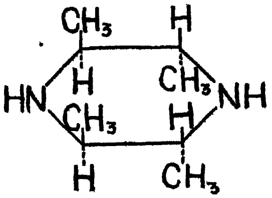
I



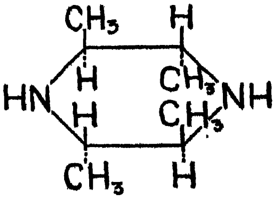
II



III



IV



V

whereas in V it was the same. Hence, if each of the nitrogens were substituted with a different group, II would give two different compounds whereas V would give one and the same compound. This could therefore be used as a means for distinguishing II from V.

In the case of the meso forms, III could be differentiated from the other two by the fact that if one of the nitrogens were substituted by another group, III would become asymmetrical and therefore resolvable.

As a result of his experiments Kipping was able to isolate four of the five possible isomers, which he designated as α, β, γ, and δ.

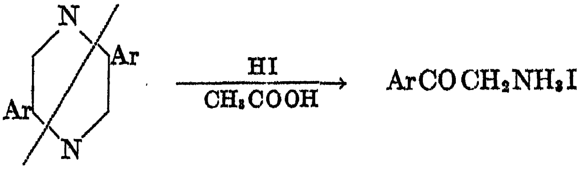
He was further able to correlate β with the structure of II, by proceeding along the lines indicated above. In the other cases structural correlation proved too difficult, because the properties of the substances were unsuitable for the operations necessary to effect the requisite separations and resolutions.

F. DEGRADATIONS OF THE NUCLEUS

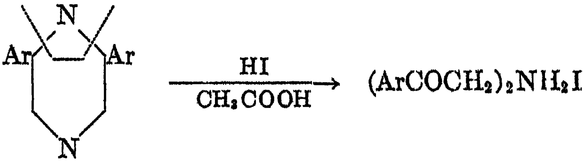
Pyrazine and its alkyl and aryl homologues are stable toward the action of acids such as sulfuric and hydrochloric, as well as alkalies. They are also quite inert to heat, since even the higher-boiling members may be distilled at atmospheric pressure without decomposition.

In contrast to pyridine, which may serve as an inert solvent for permanganate, pyrazine is susceptible to the action of this oxidizing agent (210). This is further demonstrated in most of the side-chain oxidations, where the yields are very poor because of the extensive nuclear destruction (exception: quinoxalines).

Hydrogen iodide in glacial acetic acid has been observed by Tutin (224) to cleave the nucleus of substituted pyrazines in one of two ways, depending on the orientation of the substituents. The 2,5-diarylpyrazines, as exemplified by *p,p′*-dimethoxy-2,5-diphenylpyrazine and *m,m′,p,p′*-tetramethoxy-2,5-diphenylpyrazine, were found to yield the hydroiodides of ω-amino-*p*-hydroxyacetophenone and ω-amino-*m,p*-dihydroxyacetophenone, respectively:



In the case of the 2,6-diarylpyrazines, the cleavage took another course, forming the hydroiodides of the diphenacylamines:



Thus, with 2,6-diphenylpyrazine, *p,p′*-dimethoxy-2,6-diphenylpyrazine, and *m,m′,p,p′*-tetramethoxy-2,6-diphenylpyrazine there resulted the corresponding hydroiodides of diphenacylamine, *p,p′*-dihydroxydiphenacylamine, and *m,m′,p,p′*-tetrahydroxydiphenacylamine. These two types of cleavage are interesting in that they may be viewed as being the reverse of the cyclizations leading to the formation of the pyrazines.

Gastaldi and Princivalle (91) have shown that the quaternary salts of certain hydroxypyrazines may be split with alkalies to yield the *N*-alkylamines. Specifically, methylamine, equivalent to both nitrogen atoms, was obtained when the methiodides of both 1-methyl-2,5-diphenyl- and 1-methyl-2,5-dimethyl-6-keto-1,6-dihydropyrazine were treated with alkali. This is analogous to the situation in the imidazole series, where alkalies cause the cleavage of the *N*-alkyl quaternary salts into the *N*-alkylamines (Rung–Behrend–Pinner reaction).

G. UNCLASSIFIED

Tota and Elderfield (220) cleaved the ether 2-methyl-3-ethoxyethyl-5-hydroxypyrazine with hydrogen bromide in glacial acetic acid. There resulted a compound which was believed to have the structure of 2-methyl-3,β-bromoethyl-5-hydroxypyrazine. The probability of the correctness of this structure was increased by the observation that the compound could be dehydrohalogenated with 5 per cent sodium carbonate to yield 2-methyl-3-vinyl-5-hydroxypyrazine.

In conclusion the authors wish to acknowledge the active interest of Mr. Bernard Klein during the writing of this paper.

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RECENT ADVANCES IN THE CHEMISTRY OF CARBAZOLE

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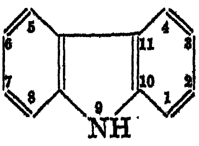
I. INTRODUCTION

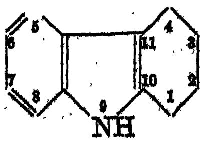
In spite of the publication of much interesting work on carbazole during the past twenty-five years, Cohn's monograph, *Die Carbazolgruppe* (1919), is the last comprehensive survey of this heterocyclic, coal-tar product. It seems appropriate, therefore, to give a brief, critical résumé of recent advances in this field. Emphasis is accordingly laid on the work of the late W. H. Perkin, Jr., and S. G. P. Plant at Oxford, S. H. Tucker at Glasgow, and Hans Lindemann at Brunswick, etc., although reference to the pioneer work of Graebe, Ullmann, Borsche, and others is naturally not neglected. No finality or completeness is claimed, our aim being to give a broad rather than a detailed picture.

II. PROPERTIES OF CARBAZOLE

In the present review, carbazole and tetrahydrocarbazole are represented by the following formulae:



Carbazole



Tetrahydrocarbazole

Various systems of numbering the carbon and nitrogen atoms are to be found in the literature. The one adopted here is that used in most American and British journals.

Carbazole is a white substance which crystallizes in beautiful flakes from hot alcohol, glacial acetic acid, benzene, or toluene. The preparation of pure carbazole is not easy, and it is not surprising that the melting points quoted in the literature vary considerably. Probably the most reliable are those of Tucker (87), 245°C.; Aristov (2), 246°C.; Senseman and Nelson (85), 244.8°C.; Zelinsky, Titz, and Gaverdowskaja (102), 245.6°C.; and Kirby (44), 247°C. It is clear from these figures that a decisive value for the melting point of carbazole is still lacking. Most samples of carbazole melt considerably below the above figures, a value of 235°C. being frequently encountered. Pure carbazole fluoresces only slightly (17, 18, 87), and it is therefore noteworthy that Aristov's sample fluoresced brilliantly.

Many color tests are used for the detection of carbazole (see Cohn's monograph, page 42). The best known is the bluish-green color observed when a trace of carbazole is dissolved in concentrated sulfuric acid and a drop of nitric acid added. Campbell and Maclean (21) found that many carbazoles—e.g., carbazole, 1-methylcarbazole, 1-aminocarbazole, 3,6-dibromocarbazole, etc.—give a dark green color with Mecke's reagent (selenious acid in concentrated sulfuric acid). The test is not specific for carbazoles, since Dewey and Gelman (26) showed that many other nitrogen compounds also give characteristic colors. Levine's statement (45) that the reagent is specific for phenolic compounds is manifestly incorrect.

Commercial carbazole gives a yellow solution in concentrated sulfuric acid, but when the carbazole is pure the solution is colorless (102).

As an indole derivative carbazole gives a red color when the vapor of its alcoholic solution comes in contact with a pine splint soaked in hydrochloric acid.

Carbazole is identified by its picrate (bright red needles, m.p. 185°C.) and trinitrobenzene derivative (red needles, m.p. 166°C.). Substituted carbazoles are similarly identified.

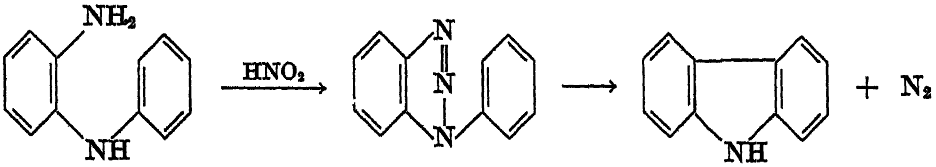
III. PREPARATION OF CARBAZOLE AND SUBSTITUTED CARBAZOLES

The main source of carbazole is coal tar, from which carbazole is obtained in large quantities. To eliminate all the impurities present, such as anthracene, is no easy task, and pure carbazole is best prepared synthetically. A few of the synthetic methods are sufficiently important to warrant some discussion here.

A. Graebe–Ullmann method

The method of Graebe and Ullmann (36) is of wide application and is exemplified by the treatment of *o*-aminodiphenylamine with nitrous acid to give 1-phenyl-1,2,3-benzotriazole, which when heated loses nitrogen to give a quantitative yield of carbazole (see page 361).

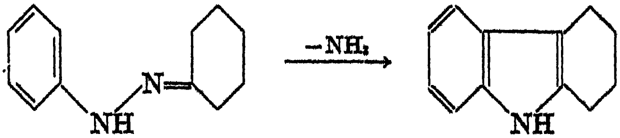
Ullmann (97) synthesized a number of carbazoles by this method, but other workers found that the presence of unsaturated groups gave negative results.



Preston, Tucker, and Cameron (93) were the first to show that nitrocarbazoles (trace), acetylcarbazoles (22 per cent), and cyanocarbazoles (34 per cent) can be obtained, although, as is clear from the yields given in parentheses, the results are not very satisfactory.

B. Borsche's method

Cyclohexanone phenylhydrazone is converted to tetrahydrocarbazole when heated with dilute sulfuric acid (5, 6, 13, 27), the process being completely analogous to the formation of indoles by the Fischer process. Borsche (13) was the first to realize the full scope of the reaction, and by it prepared many substituted tetrahydrocarbazoles. Reagents other than sulfuric acid may be used; glacial acetic acid, for example, has been found to give cleaner products (63).

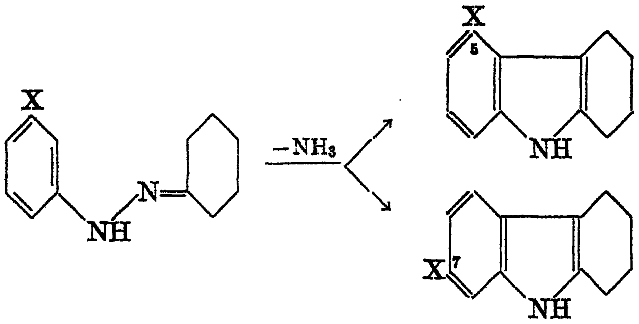


The value of the method is obviously dependent on a suitable means of effecting dehydrogenation. Many reagents have been tried, the yields obtained in most cases being small. Borsche (13) used lead oxide; Perkin and Plant (63, 64) mercurous acetate or sulfur; and Cooke and Gulland (25) palladous chloride. None of these was completely satisfactory. Barclay and Campbell (19) found chloranil to be an excellent reagent for the purpose, many substituted carbazoles being obtained in good yields. Arnold (3, 4) had previously used chloranil successfully to prepare polynuclear aromatic hydrocarbons.

An interesting dehydrogenation was accomplished by Hoshino and Takiura (39) with cinnamic acid and palladium black. The cinnamic acid was reduced to hydrocinnamic acid and carbazole was obtained in excellent yield.

Borsche's method, like that of Graebe and Ullmann, serves both for the preparation of substituted carbazoles and for the determination of their structure. Ambiguity arises, however, when the *m*-substituted phenylhydrazones of cyclohexanone are used. Two products, 5- and 7-substituted tetrahydrocarbazoles, are theoretically possible and are frequently obtained in practice (see page 362).

Plant and his collaborators have examined several reactions of this type. For example, both the 5- and the 7-carboxylic acids were obtained from *m*-hydrazinobenzoic acid (71), and their structures were later proved by Moggridge and Plant (73). Plant (69) showed also that the product from cyclohexanone *m*-nitrophenylhydrazone is a mixture of 5- and 7-nitrotetrahydrocarbazoles and not a single substance, as claimed by Borsche (13). Plant succeeded in isolating the



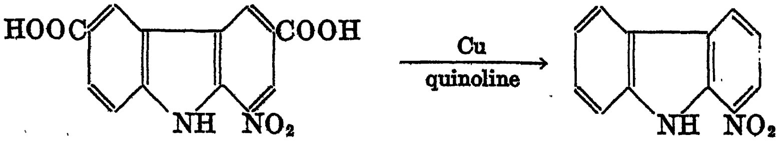
7-isomer, and Barclay and Campbell (19), using chromatographic adsorption, isolated both isomers.

C. Substitution

A limited number of substituted carbazoles can be obtained by nitration, halogenation, sulfonation, etc. but care must be exercised, since mixtures which are difficult to separate are often obtained. Mixtures containing mono-, di-, and tri-substituted isomers sometimes result. The 3- and 6-positions and to a lesser extent the 1- and 8-positions are the most reactive in carbazole. This is clearly demonstrated by nitration, which gives 75 per cent of 3-nitro- and 4 per cent of 1-nitro-carbazole (46), and by exhaustive nitration which yields 1,3,6,8-tetranitrocarbazole.

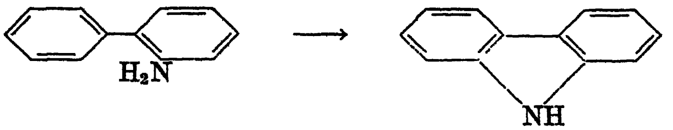
The reactivity of the carbon atoms ortho and para to the imino group in carbazole results in the formation not only of mono- but also of di-, tri-, and tetra-substituted carbazoles. By bromination, for example, 3-bromo-, 3,6-dibromo-, 1,3,6-tribromo-, and 1,3,6,8-tetrabromo-carbazoles may be prepared, depending on the quantity of bromine used. It is at once obvious that the preparation of 2- and 4-isomers by direct substitution is not possible.

Substitution is most often used to prepare the 3-substituted carbazoles. The 1-isomers are generally separated with difficulty. Morgan and Mitchell (54), for example, separated 1- and 3-nitrocarbazoles by making use of their different basicities, but Preston, Tucker, and Cameron (93) obtained more satisfactory results by chromatographic analysis. 1-Substituted carbazoles may, however, be conveniently prepared from 1,3,6-trisubstituted carbazoles by removal of the groups in the 3- and 6-positions. Tucker and coworkers (93), for example, nitrated 3,6-carbazoledicarboxylic acid to give 1-nitro-3,6-carbazoledicarboxylic acid, and then decarboxylated the acid to 1-nitrocarbazole.



D. Miscellaneous methods

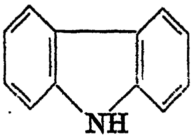
Other methods have also been used to prepare carbazole and its derivatives, many of them, however, being of little practical importance (see Cohn's monograph, page 6 *et seq.*). Mention may be made of the preparation of carbazoles from derivatives of 2,2′-diaminodiphenyl by treatment of the tetrazotized compound with copper or from *o*-xenylamine by catalytic pyrolysis. By the latter method Morgan and Walls obtained particularly pure carbazole (56).



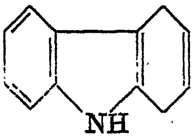
A promising method is that given in a German patent (30), according to which aromatic amines and 2-chlorocyclohexanone interact to give good yields of tetrahydrocarbazoles. The possibilities of the method are being investigated in our laboratories.

IV. REDUCTION OF CARBAZOLE

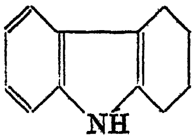
By analogy with naphthalene it might be expected that hydrogenation of carbazole would yield first 1,4-dihydro- and then 1,2,3,4-tetrahydrocarbazole.



Carbazole

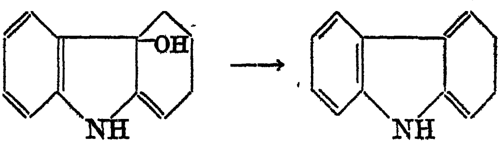


1,4-Dihydrocarbazole



Tetrahydrocarbazole

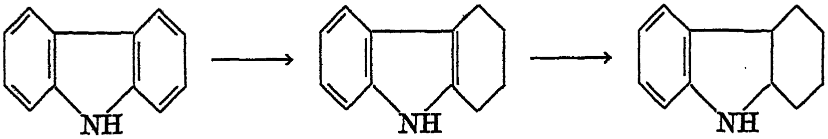
Schmidt and Schall (83) did claim to have prepared the dihydro compound by reduction of carbazole with sodium and boiling amyl alcohol, but a critical examination of the resulting so-called 1,4-dihydrocarbazole by Barclay, Campbell, and Gow (20) proved it to be a mixture containing at least 50 per cent of carbazole. This means that the only known dihydrocarbazole is 2,3-dihydrocarbazole, prepared by Plant and Tomlinson (77) by the dehydration of 11-hydroxy-2,3,4,11-tetrahydrocarbazole (see page 367).



2,3-Dihydrocarbazole

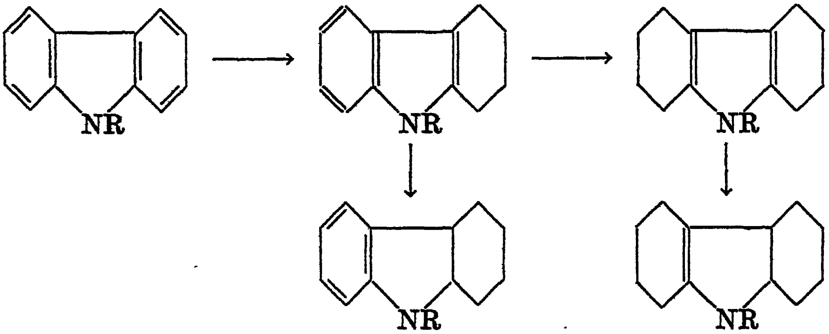
There is no ambiguity about 1,2,3,4-tetrahydrocarbazole prepared either by reduction of carbazole with sodium and alcohol (101) or more conveniently from cyclohexanone phenylhydrazone by the Borsche method (page 361). It has been stated that tetrahydrocarbazole decomposes on standing in air, but this holds only for the impure substance. When crystallized from light petroleum (100–120°C.) it separates as a colorless, stable substance, m.p. 116°C.

1,2,3,4,10,11-Hexahydrocarbazole may be obtained by reducing carbazole with hydriodic acid and phosphorus at 130°C. (84), but at higher temperatures a hydrocarbon, probably 3,3′-dimethyldicyclopentyl, is obtained. The hexahydro derivative is best prepared either by the reduction of tetrahydrocarbazole with tin and hydrochloric acid (13) or in quantitative yield by electrolytic reduction (Perkin and Plant (65)).



Tetrahydrocarbazole Hexahydrocarbazole

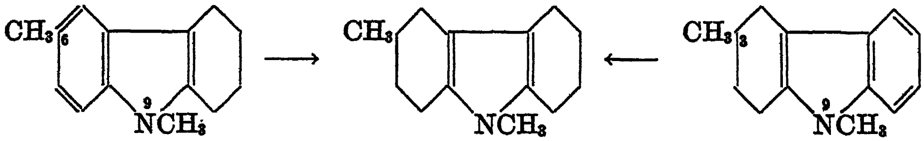
Reduction under more vigorous conditions has been investigated. von Braun and Ritter (14) found that *N*-methyl- or *N*-ethyl-carbazole with hydrogen at 210°C. and 25 atm. pressure in the presence of a nickel catalyst gives a mixture of tetra- and octa-hydrocarbazoles along with unchanged material.



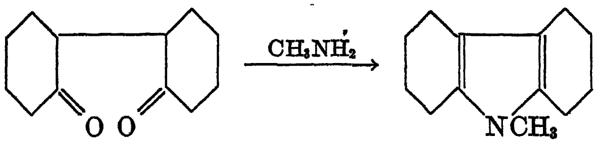
Curiously enough, the double bonds of the "pyrrole ring" in both products remain intact, but may be attacked by tin and hydrochloric acid to give hexahydro- and a decahydro-carbazole, respectively (in the form of their *N*-methyl or *N*-ethyl derivative).

The latter is extremely resistant to further reduction. Carbazole, even when purified, is not attacked under the above conditions, but this is probably due to traces of impurity.

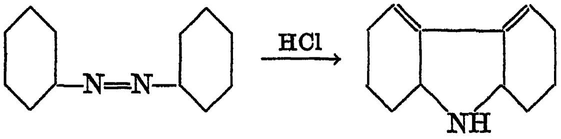
The position of the double bonds in the octahydro compound was demonstrated in two ways. von Braun and Schörnig (16) reduced 6,9-dimethyl- and 3,9-dimethyl-1,2,3,4-tetrahydrocarbazoles and obtained the same octahydro product in both cases. This result is consistent only with a pyrrole structure of the octahydrocarbazole as formulated below.



The correctness of von Braun's conclusion was confirmed by Plant's synthesis (68) from 2,2′-dicyclohexanone and methylamine—an extension of the well-known synthesis of pyrroles from 1,4-diketones.

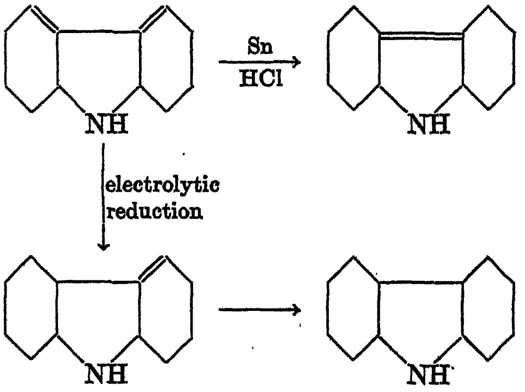


Perkin and Plant (65) prepared a second octahydrocarbazole by treating cyclohexylidine azine with hydrochloric acid.



It is highly probable that the substance has the above structure, first suggested by von Braun and Schörnig (16), but this has still to be proved. The same compound was also obtained by Benary (7) by treatment of the azine with chloroacetyl chloride at room temperature, followed by hydrolysis.

More highly reduced carbazoles have also been prepared. Perkin and Plant (65) obtained a decahydrocarbazole by reduction of their octahydrocarbazole with tin and hydrochloric acid and dodecahydrocarbazole (perhydrocarbazole) by electrolytic reduction. The decahydrocarbazole was stable to electrolytic reduction and could not therefore be an intermediate in the reduction to perhydrocarbazole. The latter substance is possibly formed as shown below:

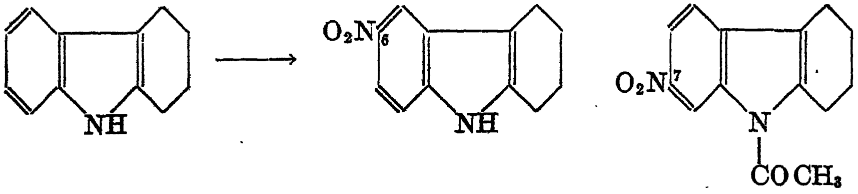


Of considerable interest are the results of Adkins and Coonvadt (1), who studied the hydrogenation of carbazole with copper chromite and Raney nickel catalysts. 1,2,3,4-Tetrahydrocarbazole, *cis*-1,2,3,4,10,11-hexahydrocarbazole, or dodecahydrocarbazole was obtained, according to the catalyst and conditions employed.

A. Properties of tetrahydrocarbazole

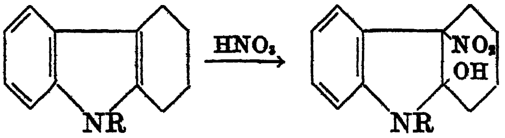
Tetrahydrocarbazole with its remarkably reactive 10,11-double bond is the most interesting of the hydrogenated carbazoles.

Nitration of tetrahydrocarbazole or its *N*-methyl derivative with concentrated nitric and sulfuric acids at −5°C. gives 6-nitrotetrahydrocarbazole or its *N*-methyl derivative, while the *N*-acetyl compound nitrates in the 7-position.

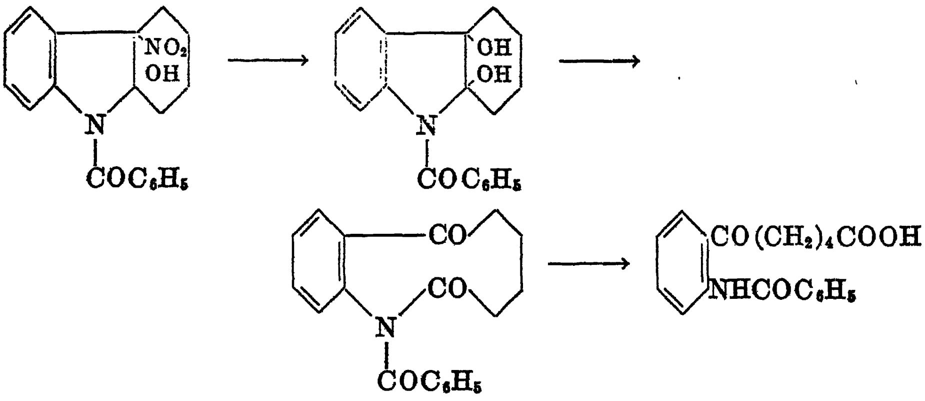


The *N*-benzoyl derivative reacts like the *N*-acetyl. Perkin and Plant (63) concluded that the *N*-benzoyltetrahydrocarbazole nitrates in the 5-position, but Plant (69) later proved this to be erroneous.

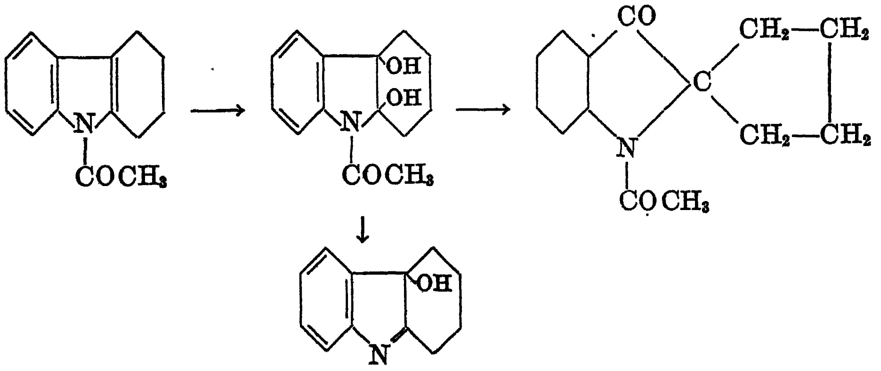
Under suitable conditions, however, *N*-substituted tetrahydrocarbazoles react with nitric acid in quite a different manner, addition of the nitric acid occurring at the double bond of the reduced ring (64). *N*-Benzoyltetrahydrocarbazole, for example, gives mainly 10-hydroxy-9-benzoyl-11-nitrohexahydrocarbazole (or, less likely, the 11-hydroxy-10-nitro derivative).



Ring fission occurs with alkali, and δ-(*o*-benzoylaminobenzoyl)valeric acid is formed. Perkin and Plant (64) pictured this reaction as taking place in the following stages.

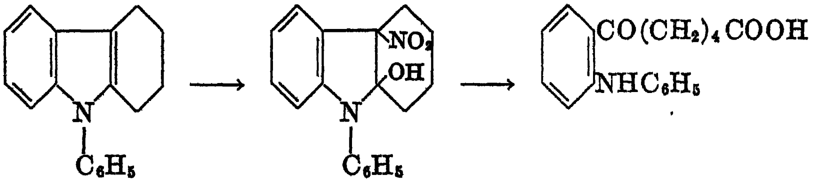


N-Acetyltetrahydrocarbazole reacts differently, 10,11-dihydroxyhexahydrocarbazole being formed. This difference is curious and difficult to explain satisfactorily. The dihydroxy derivative yields with alkali a red product, possibly 11-hydroxytetrahydrocarbazolenine, and with acetic acid a spirane, probably by a pinacol change.

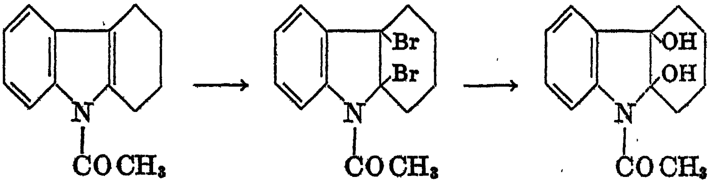


11-Hydroxytetrahydrocarbazolenine

Linnell and Perkin (62) found that N-phenyltetrahydrocarbazole behaves like the N-benzoyl compound, and 11-nitro-10-hydroxy-9-phenylhexahydrocarbazole is formed. It decomposes into *o*-anilinobenzoylvaleric acid when heated with potassium hydroxide.



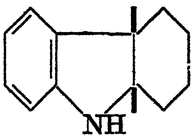
Bromination of N-substituted tetrahydrocarbazoles leads to the formation of unstable 10,11-dibromo derivatives, which react instantly with water to give the corresponding dihydroxy compounds or their anhydro derivatives (77). Thus, N-acetyltetrahydrocarbazole yields the N-acetyl derivatives of 10,11-dihydroxyhexahydrocarbazole.



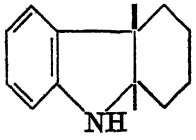
Again the benzoyl compounds behave differently from the acetyl. An intermediate dibromo product is obtained as before, but only one of the bromine atoms reacts with water, the other uniting with a neighboring hydrogen atom to form hydrobromic acid. The product is probably the benzoyl derivative of 11-hydroxy-2,3,4,11-tetrahydrocarbazole, into which it is converted on hydrolysis with alcoholic potassium hydroxide. Its conversion into 2,3-dihydrocarbazole has already been noted on page 363.

B. Properties of hexahydrocarbazole

The most interesting property of hexahydrocarbazole is its existence in two forms—*cis* and *trans*.



cis



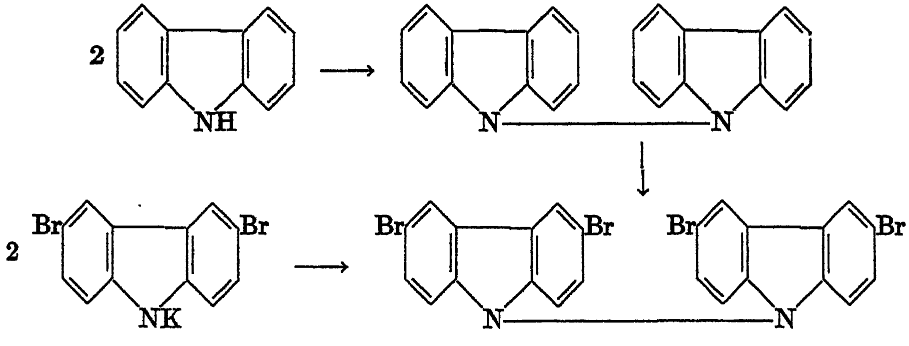
trans

Models show that the *cis* form is much less strained than the *trans*, and consequently is more stable. It is therefore to be anticipated that hydrogenation of tetrahydrocarbazole will give a great preponderance of the *cis* isomer. Gurney, Perkin, and Plant found this to be so (61), the *trans* form amounting only to 1-2 per cent of the product.

V. OXIDATION OF CARBAZOLE

Carbazoles may be oxidized by sodium dichromate in a mixture of sulfuric and acetic acids (Wieland's method (100)) or potassium permanganate in acetone. With the latter reagent Perkin and Tucker (67) isolated three products: A, m.p. 220°C.; B, m.p. 265°C.; and C (amorphous), m.p. about 175°C. A and B were shown to be dicarbazyls.

In view of the reactivity of the hydrogen atom in the 9-position it was not surprising that Tucker and McLintock (90) proved A to be 9,9′-dicarbazyl, since it gave on bromination 3,6,3′,6′-tetrabromodicarbazyl, the constitution of which followed from its preparation by the action of powdered potassium hydroxide followed by iodine on 3,6-dibromocarbazole.



It was expected that the product B would be 3,3′-dicarbazyl, but Tucker (88) synthesized this compound and showed it to be different from B. 3,3′-Dicarbazyl was, however, obtained by the oxidation of carbazole with Wieland's reagent.

The possibility that B might be 1,1′- or 3,9′-dicarbazyl was ruled out by Tucker's synthesis of these substances (91, 92). Their properties differed from those of B, which is presumably 1,3′- or 1,9′-dicarbazyl.

VI. SIMPLE DERIVATIVES OF CARBAZOLE

It is only in recent times that many of the simple carbazole derivatives have been prepared, and a short account of them is therefore in order.

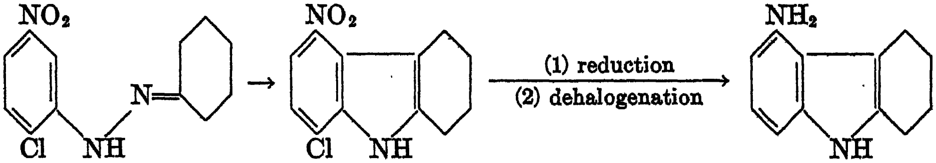
A. *Nitrocarbazoles*

3-Nitrocarbazole (m.p. 214°C.) is the most accessible of the nitrocarbazoles and is best prepared by modifications of the method of Ruff and Stein (43, 46, 82, 94), by the action of nitric acid on *N*-nitrosocarbazole. The nitroso group is then readily removed by heating with alkali or glacial acetic acid.

Two other mononitrocarbazoles have been reported in the literature. One, m.p. 187°C., was first isolated by Votocek (98) in 1896. Lindemann (46) showed that it is obtained during the preparation of 3-nitrocarbazole in about 4 per cent yield, and proved it to be 1-nitrocarbazole by reducing it to 1-aminocarbazole, identical with a sample synthesized by the Graebe–Ullmann method (48, 49). The other nitro compound, m.p. 164°C., was obtained by Ziersch (103) and was for long mistaken for 1-nitrocarbazole, but Morgan and Mitchell (54) showed it to be a molecular compound of 1- and 3-nitrocarbazoles. Failure to realize that it was a mixture has led to some confusion (see, for example, Whitner (99)).

1-Nitrocarbazole may be prepared from 1-nitro-3,6-carbazoledicarboxylic acid (see page 362) or by removal of the sulfonyl groups from 1-nitro-3,6,8-carbazoletrisulfonic acid (58).

Borsche, Witte, and Bothe (13) obtained a supposedly single product from the ring-closure of cyclohexanone *m*-nitrophenylhydrazone, but Plant (69) showed that in reality it was a mixture of 5- and 7-nitrotetrahydrocarbazoles, although a pure sample of the 5-isomer was not obtained. The structures of these compounds were settled by reduction to the corresponding amines and comparison of their properties with a sample of 5-aminotetrahydrocarbazole, the constitution of which follows from its synthesis from cyclohexanone 2′-chloro-5′-nitrophenylhydrazone.



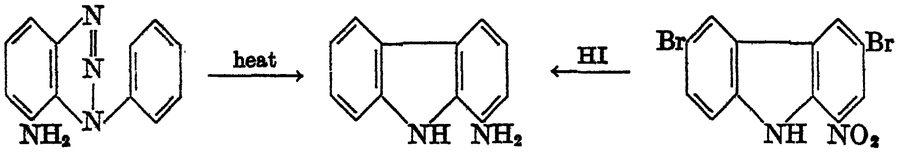
Barclay and Campbell (19) succeeded in obtaining pure samples of both the 5- and the 7-nitrotetrahydrocarbazoles by chromatographic separation, and prepared 2- and 4-nitrocarbazoles from them by dehydrogenation with chloranil.

B. *Aminocarbazoles*

Our knowledge of the amines is, of course, dependent on the isolation of the corresponding nitrocarbazoles. It follows that 3-aminocarbazole, readily prepared in quantity by the reduction of 3-nitrocarbazole (43, 99), has been thoroughly studied. Of the other isomers, 1-aminocarbazole has been prepared, 2-aminocarbazole *may* have been obtained, and 4-aminocarbazole has not been described in the literature.

Lindemann and Werther (48) were the first to prepare pure 1-aminocarbazole. This was done both by the reduction of 1-nitrocarbazole and by synthesis from

1-phenyl-7-aminobenzotriazole, but neither method gave large yields. Better results were obtained from 1-phenyl-5-carboxy-7-aminobenzotriazole (Lindemann and Wessel (49)), but this method is also unsuitable for large-scale preparations. A more promising method is that of Campbell and Maclean (21), in which 3,6-dibromo-1-nitrocarbazole is treated with hydriodic acid. Reduction of the nitro group is accompanied by removal of the bromine atoms. The method suffers from the disadvantage that the resulting 1-aminocarbazole is sometimes difficult to purify. Better results might be obtained by the use of a milder reagent, such as Raney nickel.



1-Aminocarbazole

A product claimed by Whitner (99) to be 1-aminocarbazole was probably impure 3-aminocarbazole.

2-Aminocarbazole is stated in the literature to be a substance, m.p. 238°C., isolated by Blank (10) from the pyrolysis of "diphenylin" (presumably 2,4'-diaminodiphenyl). Blank's claims are of doubtful validity, as his paper contains several unsatisfactory features, such as lack of details regarding his starting material and no analytical figures for the resulting aminocarbazole.

C. Halogenocarbazoles

Halogenation of carbazole proceeds as would be expected, 3-halogenocarbazoles being first formed. These, however, readily react with more halogen to give 3,6-dihalogenocarbazoles, and these are the most accessible halogen derivatives of carbazole. The isolation of 3-carbazolediazonium chloride in a crystalline condition by Morgan and Read (55) opened up another route for the preparation of halogenocarbazoles, and Tucker (89) used it for this purpose.

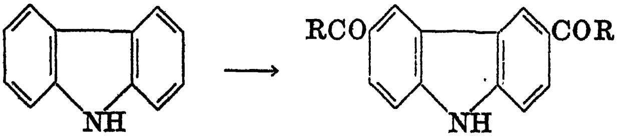
All the monochlorocarbazoles have been prepared. Moggridge and Plant (73) obtained 5- and 7-chlorotetrahydrocarbazoles (the former impure) from cyclohexanone *m*-chlorophenylhydrazone, and prepared 2- and 4-chlorocarbazoles from them by dehydrogenation with sulfur in quinoline. Barclay and Campbell (19) obtained all the chlorocarbazoles by chloranil dehydrogenation of the corresponding chlorotetrahydrocarbazoles. The four bromocarbazoles have been prepared by this method.

Tucker (87) was the first to prepare pure iodocarbazoles by the action of potassium iodide and potassium iodate on carbazole in glacial acetic acid. The main product was 3-iodocarbazole, identical with a specimen prepared by the diazotization of 3-aminocarbazole. Tucker also prepared 3,6-diiodocarbazole and showed that Classen's diiodocarbazole (24) was an impure product. The positions of the iodo groups are not proved, but are almost certainly 3 and 6 by analogy with other disubstitution products of carbazole. Moreover, Gilman and Kirby (33) showed that *N*-ethylcarbazole on iodination gave the 3,6-diiodo-*N*-ethylcarbazole.

D. Carbazyl ketones

The acylcarbazoles are prepared by the Friedel–Crafts reaction and have been thoroughly studied by Plant and his coworkers. The course of the acylation depends on several factors (see below).

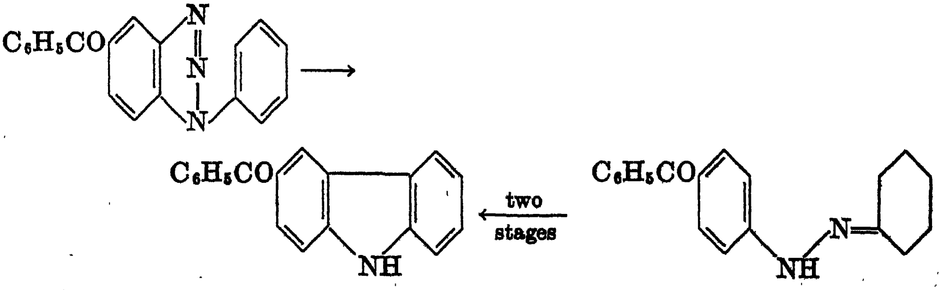
Carbazole with aluminum chloride and acid chlorides (even in limited quantities) in carbon disulfide yields 3,6-disubstituted products, monosubstituted derivatives being obtained only in limited quantities. In this way Plant (76, 78) obtained 3,6-dibenzoyl- and 3,6-diacetyl-carbazoles.



R = CH_3 or C_6H_5.

The structures of these compounds were determined by comparison with samples synthesized by unambiguous methods.

3-Benzoylcarbazole was prepared from 5-benzoyl-1-phenylbenzotriazole by Hunter and Darling (42) and from *p*-hydrazinobenzophenone and cyclohexanone by Plant and Tomlinson (78). Both syntheses proved the structure of the product.



Attempts to prepare 1-benzoylcarbazole from 7-benzoyl-1-phenyltriazole failed.

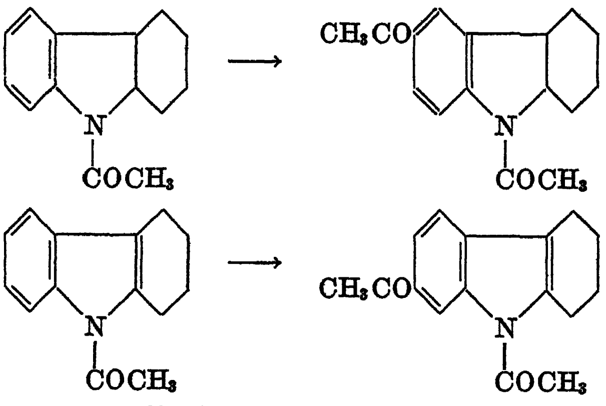
3-Benzoylcarbazole was also obtained by "baking" (Scholl process) *N*-benzoylcarbazole with aluminum chloride at 120°C. (78), and the 3-acetyl compound was obtained in an analogous manner (79). The constitution of the latter substance was proved by reduction to 3-ethylcarbazole, whose structure was determined by synthesis. Plant (76) and Meitzner (52) obtained the best results with nitrobenzene as solvent, and Meitzner succeeded in isolating some 1-acetylcarbazole in addition to the main product, 3-acetylcarbazole (50–60 per cent).

9-Alkylcarbazoles react with acyl chlorides and aluminum chloride in the same way as carbazole, 9-methylcarbazole, for example, giving with acetyl bromide and aluminum chloride 3,6-diacetyl-9-methylcarbazole (76). 9-Acylcarbazoles, on the other hand, give 2,9-derivatives. Plant and Williams

(79) found that 9-acetylcarbazole gave 2,9-diacetylcarbazole (*cf.* Borsche and Feise (12)). Results with 9-benzoylcarbazole have not been consistent. Plant and Tomlinson (78) found that it yielded with benzoyl chloride and aluminum chloride 3,6-dibenzoylcarbazole, but later Plant, Williams, and Rogers (76) were unable to repeat this result and obtained instead 2,9-dibenzoylcarbazole. These discordant observations have not been completely explained. The reactivity of the 2-position in the above experiments was also shown by Ruberg and Small (81), who obtained 2-chloroacetyl-9-acetylcarbazole from 9-acetylcarbazole and chloroacetyl chloride.

Hydrolysis of the 2,9-derivatives gives 2-acetyl- and 2-benzoylcarbazoles. Plant, Rogers, and Williams (76) obtained 2-benzoylcarbazole (m.p. 163°C.), different from the substance (m.p. 350°C.) previously regarded as having this structure (31).

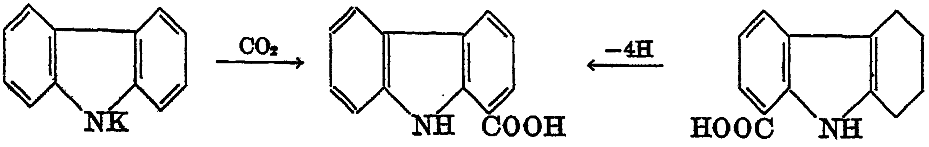
Another factor affecting the course of the Friedel–Crafts reaction is the degree of reduction of the carbazole. Mitchell and Plant (72) obtained 6,9-diacetylhexahydrocarbazole from 9-acetylhexahydrocarbazole, but Plant and Rogers (75) found the product from 9-acetyltetrahydrocarbazole to be 7,9-diacetyltetrahydrocarbazole.



E. Carbazolecarboxylic acids

The carbazolecarboxylic acids are not readily prepared in large quantities, but 1-, 2-, and 3-carbazolecarboxylic acids have been prepared and characterized.

Ciamician and Silber (23) obtained a monocarboxylic acid, m.p. 271–272°C., by passing carbon dioxide into potassium carbazole, followed by treatment with dilute sulfuric acid. The assumption that it was the 1-carboxylic acid was proved many years later, when Briscoe and Plant (70) showed it to be identical with the dehydrogenation product of 1,2,3,4-tetrahydro-8-carbazolecarboxylic acid.

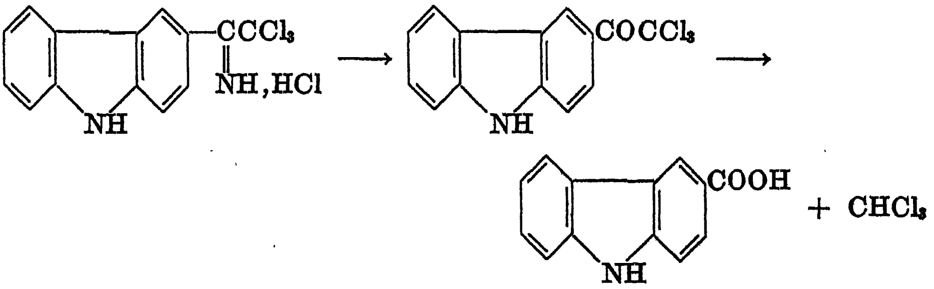


Gilman and Kirby (33) found that interaction of carbazole and *n*-butyllithium followed by carbonation gave 1-carbazolecarboxylic acid in very poor yield. Better results were obtained with *N*-ethylcarbazole, the resulting 9-ethyl-1-carbazolecarboxylic acid being identical with that obtained by treating 1-carbazolecarboxylic acid with diethyl sulfate and potassium hydroxide.

2-Carbazolecarboxylic acid was prepared by Borsche and Feise (12) by fusing 2-acetylcarbazole with potassium hydroxide (*cf.* Plant and Williams (79)); by Moggridge and Plant (73) by dehydrogenation of 1,2,3,4-tetrahydro-7-carbazolecarboxylic acid with palladized charcoal; and from 2-benzoylcarbazole by fusion with potassium hydroxide (76).

3-Carbazolecarboxylic acid has been prepared by similar methods: fusion of 3-acetylcarbazole with potassium hydroxide (79); dehydrogenation of the methyl ester of the corresponding tetrahydrocarbazolecarboxylic acid (73); and mercuration of *N*-ethylcarbazole and suitable treatment of the product (33), 9-ethyl-3-carbazolecarboxylic acid being obtained. It is of interest that whereas metallation of *N*-ethylcarbazole with lithium gives a 1-lithium compound, mercuration occurs in the 3-position.

More recently Dunlop and Tucker (89) have prepared the acid by the Houben–Fischer method (40, 41). Trichloroacetonitrile acted on carbazole in chlorobenzene solution with aluminum chloride and dry hydrogen chloride, followed by hydrolysis, to give 3-trichloroacetylcarbazole. Alkaline hydrolysis split off chloroform and yielded 3-carbazolecarboxylic acid.



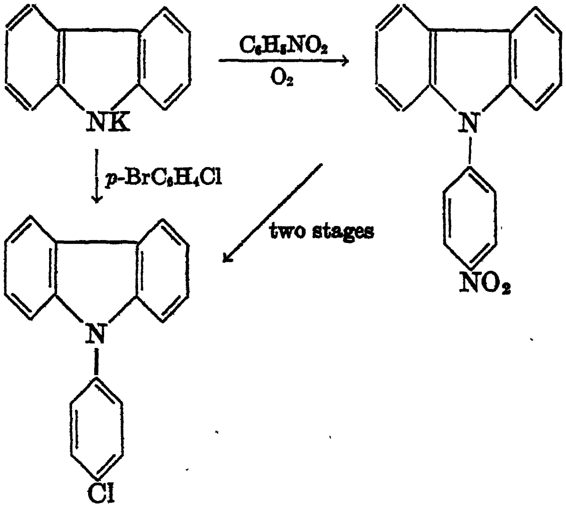
4-Carbazolecarboxylic acid has not yet been prepared, since efforts to dehydrogenate 1,2,3,4-tetrahydro-5-carbazolecarboxylic acid were not successful (73).

A carbazoledicarboxylic acid is stated in the patent literature to have been prepared by the action of carbon dioxide on potassium carbazole (29), but no proof of structure is given. Mitchell and Plant (72) were the first to prepare 3,6-carbazoledicarboxylic acid and prove its structure by showing it to be identical with the product of fusion of 3,6-dibenzoylcarbazole with potassium hydroxide. It has also been prepared by the Houben–Fischer method (89).

F. N-Substituted carbazoles

N-Alkyl, -aryl, or -acyl derivatives have been obtained by the action of the appropriate reagents on (*1*) carbazole, (*2*) potassium carbazole, (*3*) or carbazole-magnesium iodide. For example, Boeseken (11) prepared *N*-acetylcarbazole in good yield by the action of acetic anhydride containing a trace of sulfuric acid on carbazole. None of these methods is completely satisfactory, and it was therefore a considerable advance when Tucker and Stevens (94) showed that nearly quantitative yields of *N*-derivatives are obtained by the action at room temperature of the appropriate alkylating or acylating reagent upon carbazole in acetone or alcohol solution in the presence of sodium or potassium hydroxide. These authors also found that the greater the acidity of the imino hydrogen, the more successful is the method. The process has occasionally proved to be capricious or ineffective. Tucker (87) found that vigorous shaking must be avoided in the preparation of *N*-benzoylcarbazole, and Ruberg and Small (81) were able to methylate 2-chloroacetylcarbazole only to a slight extent. Nevertheless the method is of wide application. It may be noted that the *N*-benzoyl compound is obtained in 60 per cent yield by heating carbazole with anhydrous potassium carbonate and benzoyl chloride in the presence of copper bronze (89).

An interesting reaction is the oxidative condensation of *N*-potassium carbazole with nitrobenzene (53). It was shown by de Montmollin and de Montmollin (53) that condensation occurred in the para position of the nitrobenzene. The 9-*p*-nitrophenylcarbazole on reduction, followed by diazotization, etc., gave a product identical with that prepared from *N*-potassium carbazole and *p*-chlorobromobenzene.



This was confirmed by the synthesis of the nitro compound by Nelmes and Tucker (92) from carbazole and *p*-chloronitrobenzene in the presence of potassium carbonate.

Patterson and Adams (59) studied other *N*-phenylcarbazole derivatives and showed that while *N*-*o*-carboxyphenylcarbazole could not be resolved, *N*-*o*-carboxyphenyl-3-nitrocarbazole (see page 376) could be resolved. The lack of symmetry in this case is clearly due to restricted rotation.

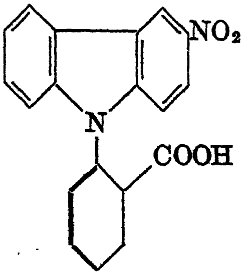
TABLE 1

Simple derivatives of carbazole

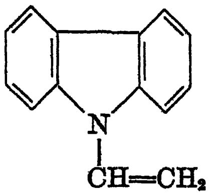
(For numbering of ring, see page 359)

COMPOUND	MELTING POINT °C.	REFERENCES
1-Nitrocarbazole	187	(46, 93)
	185–187	(19)
2-Nitrocarbazole	165–166	(19)
3-Nitrocarbazole	214	(46)
	209–210	(82)
	203–206	(19)
	No. m.p.	(94)
4-Nitrocarbazole	182–183	(19)
3,6-Dinitrocarbazole	320	(103)
1-Aminocarbazole	193	(21, 48)
2-Aminocarbazole	238	(10)
3-Aminocarbazole	254	(99)
	259	(82)
1-Chlorocarbazole	125	(58)
	109–110	(19)
2-Chlorocarbazole	244	(97)
	242	(19, 73)
3-Chlorocarbazole	201.5	(97)
	195	(86)
	199–200	(19)
4-Chlorocarbazole	96	(19, 73)
1,4-Dichlorocarbazole	84–85	(19)
3,6-Dichlorocarbazole	202–203	(51)
1-Bromocarbazole	111–112	(19)
2-Bromocarbazole	250–251	(19)
3-Bromocarbazole	197	(64)
	194–195	(86)
	201–202	(19)
4-Bromocarbazole	104–105	(19)
3,6-Dibromocarbazole	213	(47)
3,6-Dibromo-1-nitrocarbazole	260	(47)
3,6-Dibromo-1-aminocarbazole	192	(47)
3-Iodocarbazole	192–194	(86, 87)
3,6-Diiodocarbazole	202–204	(87)
1-Acetylcarbazole	136	(52)
2-Acetylcarbazole	227	(12, 79)
3-Acetylcarbazole	167	(79)
	162	(76)
	165	(93)
3,6-Diacetylcarbazole	232	(76)
2-Benzoylcarbazole	163	(76)
3-Benzoylcarbazole	203–205	(42)
	206	(78)
3,6-Dibenzoylcarbazole	258	(78)
3-Ethoxycarbazole	106–107	(39)
	105–106	(19)

COMPOUND	MELTING POINT	REFERENCES
	°C.	
2-Ethylcarbazole	225	(79)
3-Ethylcarbazole	144	(79)
3,6-Diethylcarbazole	119	(76)
1-Methylcarbazole	120.5	(97)
	110–114	(19)
2-Methylcarbazole	259	(13, 19)
3-Methylcarbazole	207	(96, 74)
	203	(13, 97)
	199–202	(19)
1-Carbazolecarboxylic acid	270–271	(70, 73)
	271–273	(19)
2-Carbazolecarboxylic acid	320–322	(76, 79)
3-Carbazolecarboxylic acid	276–278	(73, 79)
	272–274	(19)
9-Acetylcarbazole	69	(35)
	75	(64)
	76	(11)
	No m.p.	(8)
9-Benzoylcarbazole	95.5	(9)
	98.5	(50)
	98	(87, 94)
	No m.p.	(89)
9-Ethylcarbazole	70	(95)
	No m.p.	(94)
9-Methylcarbazole	85–86	(63, 64)
	87	(34)
	88	(95)
	No m.p.	(94)
9-Phenylcarbazole	82–84	(22)
	88–89	(19, 38)
	91–93	(89)
	94–95	(28)
3,9-Diacetylcarbazole	153	(79)
3,9-Dibenzoylcarbazole	170	(78)
9-Acetyl-3-benzoylcarbazole	154	(78)
9-Benzoyl-2-acetylcarbazole	153	(79)
9-Methyl-2-acetylcarbazole	122	(79)
9-Methyl-3-acetylcarbazole	102	(79)



N-*o*-Carboxyphenyl-3-nitrocarbazole



N-Vinylcarbazole

TABLE 2

Simple derivatives of tetrahydro- and hexahydro-carbazoles

(For numbering of ring, see page 359)

COMPOUND	MELTING POINT °C.	REFERENCES
Tetrahydrocarbazole	116	(13)
	119	(63)
Hexahydrocarbazole	96	(13)
	99	(65)
5-Nitrotetrahydrocarbazole	155–156	(19)
6-Nitrotetrahydrocarbazole	174	(13, 63)
7-Nitrotetrahydrocarbazole	172	(19, 69)
8-Nitrotetrahydrocarbazole	148–149	(13)
	149–150	(66)
7-Nitrohexahydrocarbazole	69	(69)
5-Aminotetrahydrocarbazole	163	(69)
6-Aminotetrahydrocarbazole	152	(63)
7-Aminotetrahydrocarbazole	101	(69)
8-Aminotetrahydrocarbazole	163–164	(66)
5-Chlorotetrahydrocarbazole	Syrup	(19, 73)
6-Chlorotetrahydrocarbazole	138	(13, 19)
7-Chlorotetrahydrocarbazole	181	(19, 73)
8-Chlorotetrahydrocarbazole	55–56	(13, 19)
5,8-Dichlorotetrahydrocarbazole	91–93	(19)
5-Bromotetrahydrocarbazole	Syrup	(80)
	131–132	(19)
6-Bromotetrahydrocarbazole	153	(13, 19)
7-Bromotetrahydrocarbazole	183	(80)
	171–172	(19)
8-Bromotetrahydrocarbazole	Syrup	(19)
1-Methyltetrahydrocarbazole	72	(37)
2-Methyltetrahydrocarbazole	98–100	(13)
	94	(19)
6-Methyltetrahydrocarbazole	141–142	(13, 19)
8-Methyltetrahydrocarbazole	97–98	(19)
6-Ethoxytetrahydrocarbazole	87–88	(13)
	105–106	(39)
	102–104	(19)
3-Carboxytetrahydrocarbazole	195	(60)
5-Carboxytetrahydrocarbazole	210	(57, 71, 73)
6-Carboxytetrahydrocarbazole	282	(57, 71)
7-Carboxytetrahydrocarbazole	287	(57, 71, 73)
8-Carboxytetrahydrocarbazole	203	(57, 71)
9-Acetyltetrahydrocarbazole	77	(63)
9-Benzoyltetrahydrocarbazole	85	(64)
9-Methyltetrahydrocarbazole	50	(63, 94)
9-Phenyltetrahydrocarbazole	86	(62)

Finally, *N*-vinylcarbazole (see page 376) may be mentioned. It is prepared in a number of ways, including the interaction of carbazole and acetylene with an alkali catalyst under pressure and at moderate temperatures (32). It readily yields a transparent thermoplastic polymer with the very high softening temperature of 250°C. It comes on the market as Luvican.

VII. TABLES OF DERIVATIVES

Table 1 lists the names and melting points of simple derivatives of carbazole, while table 2 gives the names and melting points of simple derivatives of reduced carbazoles.

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REACTIONS OF THE HALOGENS WITH THE SILVER SALTS OF CARBOXYLIC ACIDS

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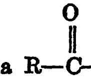
The nature of the product of reaction between silver salt and halogen is dependent upon five factors: (*a*) the nature of the silver salt; (*b*) the ratio of silver salt to halogen; (*c*) the temperature of reaction; (*d*) the presence or absence of other reactive materials; and (*e*) the effect of the solvent. Reaction between equivalent quantities of the silver salt of a carboxylic acid and halogen in an inert solvent produces a complex intermediate, which may be thermally decomposed to give the ester of the corresponding acid and the next lower alcohol. In benzene solutions of olefinic substances this complex intermediate reacts to give the dicarboxylate. Reaction between one equivalent of a silver salt of a carboxylic acid and one mole of halogen produces an active intermediate of the formula  $R{-}\overset{\overset{O}{\|}}{C}{-}OX$, in which the halogen apparently has positive character. This substance adds rapidly to olefins, is decomposed thermally to give, by decarboxylation, the corresponding halides, and may be stabilized by coördination of the positive halogen (provided it is iodine) with pyridine.

Descriptions of reactions of the halogens with silver salts of carboxylic acids have appeared periodically in the chemical literature for many years. Peligot (20), as early as 1836, described a reaction between silver benzoate and bromine in which *m*-bromobenzoic acid was isolated. A critical survey of the literature of this subject shows that the nature of the product of reaction between silver salt and halogen is dependent upon five factors: (*a*) nature of the silver salt; (*b*) ratio of silver salt to halogen; (*c*) temperature of reaction; (*d*) presence or absence of other reactive materials; and (*e*) effect of solvent.

The reactions fit definite patterns only when the ratio of equivalents of silver salt to halogen is 1:2 or 1:1. Therefore, cases involving ratios other than these are excluded from this report.

I. REACTIONS INVOLVING EQUIVALENT QUANTITIES OF SILVER SALT AND HALOGEN (SEE TABLE 1)

A. The action of salts of monocarboxylic acids

Work first performed by Simonini (30) and substantiated by other investigators (12, 22, 25, 32) shows that equivalent quantities of iodine and silver salts of monocarboxylic acids react at moderate temperatures to give an active complex intermediate which can be isolated.

$$2RCOOAg + I_2 \xrightarrow[\substack{\text{or}\\ \text{benzene}}]{\text{ether}} (RCOO)_2AgI + AgI$$

TABLE 1

Summary of specific reactions for equivalent quantities of silver salt and halogen

SILVER SALT OF ACID	HALOGEN	REMARKS	ORGANIC PRODUCTS	REFERENCES	YIELD
Acetic acid	Iodine	Solid state; *ca.* 100°C.	Methyl acetate and acetic acid	(2, 11, 29)	50–70 per cent of ester
Acetic acid	Iodine	Petroleum ether as solvent	Complex intermediate	(30, 32)	
Phenylacetic acid	Iodine	Solid state; *ca.* 80°C.	Benzyl phenylacetate and phenylacetic acid	(32)	68 per cent of ester
Triphenylacetic acid	Iodine	Benzene as solvent; 80°C.	Triphenylmethyl ester of triphenylacetic acid	(32)	Quantitative
Caproic acid	Iodine	Solid state; *ca.* 100°C.	Chiefly amyl caproate	(29, 32)	70 per cent
Caproic acid	Iodine	Petroleum ether as solvent	Complex intermediate	(30)	
Palmitic acid	Iodine	Solid state; *ca.* 100°C.	Pentadecyl palmitate	(8, 19)	
Stearic acid	Iodine	Solid state; *ca.* 100°C.	Heptadecyl stearate	(8, 11)	
Lacceric acid	Iodine	Solid state; *ca.* 100°C.	Myricyl laccerate	(9)	
Benzoic acid	Iodine	Ethyl ether or benzene as solvent	Complex intermediate	(12, 22, 32)	
Benzoic acid	Iodine	Decomposition of complex; 150–190°C.	Iodobenzene, benzoic acid, and phenyl benzoate	(32)	Small amount of ester
Glutaric acid	Iodine	Solid state; *ca.* 90°C.	γ-Butyrolactone and glutaric anhydride	(32, 33)	30 per cent of lactone
Adipic acid	Iodine	Solid state; *ca.* 90°C.	δ-Valerolactone and adipic acid	(32)	Low yield of lactone
β,β-Dimethylglutaric acid	Iodine	Solid state; *ca.* 90°C.	β,β-Dimethyl-γ-butyrolactone and β,β-dimethylglutaric anhydride	(33)	40 per cent of lactone
α-Ethylglutaric acid	Iodine	Solid state; *ca.* 90°C.	γ-Caprolactone	(33)	
α,α'-Diethylglutaric acid	Iodine	Solid state; *ca.* 90°C.	α-Ethyl-γ-caprolactone	(33)	
1,1-Cyclohexanediacetic acid	Iodine	Solid state; *ca.* 90°C.	Lactone of 1,1-cyclohexanediacetic acid	(34)	
4-Methyl-1,1-cyclohexanediacetic acid	Iodine	Solid state; *ca.* 100°C.	Lactone of 4-methylcyclohexane-1,1-diacetic acid	(10)	16 per cent

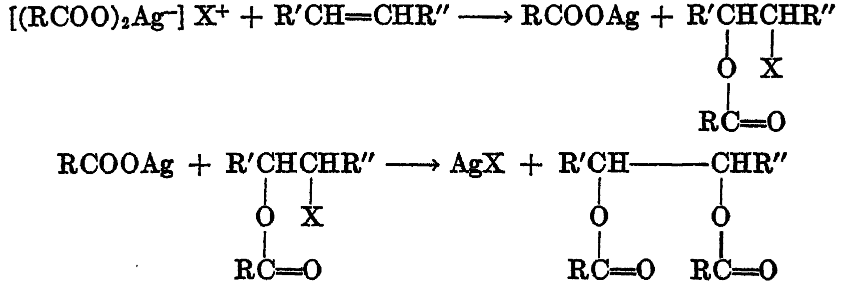
is-Hexahydrohomophthalic acid	Iodine	Solid state; *ca.* 90°C.	*cis*-Hexahydrophthalide and *cis*-hexahydrohomophthalic acid	(34)	
rans-Hexahydrohomophthalic acid	Iodine	Solid state; *ca.* 90°C.	*cis*-Hexahydrophthalide and *trans*-hexahydrohomophthalic acid	(34)	
is-Hexahydroisophthalic acid	Iodine	Solid state; *ca.* 90°C.	Lactone of *cis*-hexahydroisophthalic acid	(34)	
;amphoric acid	Iodine	Solid state; *ca.* 90°C.	Anhydride of camphoric acid	(34)	
’hthalic acid	Iodine	Solid state; *ca.* 130°C.	Phthalic anhydride	(32)	84 per cent
;lycolic acid	Iodine	Solid state	Formaldehyde and glycolic acid	(32)	
,actic acid	Iodine	Benzene as solvent; room temperature	Acetaldehyde and lactic acid	(32)	
Iandelic acid	Iodine	Benzene as solvent; room temperature	Benzaldehyde and mandelic acid	(32)	60 per cent of aldehyde
Benzilic acid	Iodine	Room temperature	Complex intermediate	(32)	
Benzilic acid	Iodine	Decomposition of complex; 55–60°C.	Benzophenone and benzilic acid	(32)	42 per cent of acid
\cetic acid	Iodine	Benzene solution of 1-pentene	1,2-Pentanediol diacetate	(23)	Good
\cetic acid	Iodine	Benzene solution of unsymmetrical diphenylethylene	Glycol diacetate of unsymmetrical diphenylethylene	(23)	Good
Benzoic acid	Iodine	Benzene solution of ethylene	1,2-Ethanediol dibenzoate	(23)	Good
Benzoic acid	Iodine	Benzene solution of propylene	1,2-Propanediol dibenzoate	(23)	Good
Benzoic acid	Iodine	Benzene solution of 1-pentene	1,2-Pentanediol dibenzoate	(23)	Good
Benzoic acid	Iodine	Benzene solution of 1-hexadecene; hydrolysis of product	1,2-Hexadecanediol	(18)	33 per cent
Benzoic acid	Iodine	Benzene solution of 1-eicosene; hydrolysis of product	1,2-Eicosanediol	(18)	70 per cent
Benzoic acid	Iodine	Benzene solution of 1-octadecene; hydrolysis of product	1,2-Octadecanediol	(18)	73 per cent
Benzoic acid	Chlorine, bromine, or iodine	Carbon tetrachloride solution of olefinic compounds; no further details	Isolation of the benzoates of the corresponding halo alcohols	(25, 28)	
Benzoic acid	Iodine	Benzene solution of styrene	Glycol dibenzoate of styrene	(23)	Good

TABLE 1—*Concluded*

SILVER SALT OF ACID	HALOGEN	REMARKS	ORGANIC PRODUCTS	REFERENCES	YIELD
Benzoic acid	Iodine	Benzene solution of unsymmetrical diphenylethylene	Glycol dibenzoate of unsymmetrical diphenylethylene	(23)	Good
Benzoic acid	Iodine	Benzene solution of stilbene	Glycol dibenzoate of stilbene	(23)	Good
Benzoic acid	Iodine	In presence of allylbenzene; hydrolysis of product	3-Phenyl-1,2-propanediol	(12)	Good
Benzoic acid	Iodine	Benzene solution of $RCH{=}CHCH_2OCOC_6H_5$; no further details	Dibenzoate of corresponding glycol	(22)	70 per cent
Benzoic acid	Iodine	Benzene solution of $RCH{=}CHOCOC_6H_5$; no further details	Dibenzoate of corresponding glycol	(22)	35 per cent
Benzoic acid	Iodine	Benzene solution containing excess butadiene; hydrolysis of products	80 per cent of 1,2-dihydroxy-3-butene and 4 per cent of 1,4-dihydroxy-2-butene	(26)	←
Benzoic acid	Iodine	Benzene solution of acetylene	Iodo- and diiodo-acetylenes	(24)	
Benzoic acid	Chlorine, bromine, or iodine	Benzene solution of 1-heptyne	Corresponding haloheptyne	(28)	Good
Benzoic acid	Iodine	Benzene solution of phenylacetylene	1-Iodo-2-phenylacetylene	(24)	

Upon hydrolysis the complex yields the parent acid, silver iodide, and silver iodate, indicating that the iodine in the complex has positive character and has undergone an autoöxidation–reduction. Thermal decomposition of the intermediate gives, as the main product, the ester of the corresponding acid and the next lower alcohol. The ester, along with some free acid, is also formed at elevated temperatures in solid-state reactions between equivalent quantities of halogen.

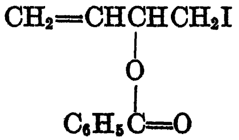
Prévost (22) first demonstrated that in the presence of benzene solutions of olefinic substances, the complex intermediate, described above, reacts to give the dicarboxylate, which upon hydrolysis yields the corresponding glycol in good yield. With olefinic materials capable of *cis–trans* isomerism, the *trans* addition predominates (23). According to Prévost (25, 27), the reaction proceeds in two steps:



$$[(RCOO)_2Ag^-]\,X^+ + R'CH{=}CHR'' \longrightarrow RCOOAg + R'\underset{\substack{|\\ O\\ |\\ RC{=}O}}{C}H\underset{\substack{|\\ X}}{C}HR''$$

$$RCOOAg + R'\underset{\substack{|\\ O\\ |\\ RC{=}O}}{C}H\underset{\substack{|\\ X}}{C}HR'' \longrightarrow AgX + R'\underset{\substack{|\\ O\\ |\\ RC{=}O}}{C}H{-\!-\!-}\underset{\substack{|\\ O\\ |\\ RC{=}O}}{C}HR''$$

When carbon tetrachloride is used as solvent (25, 28) the halogen esters can be isolated, since the second phase of the reaction takes place very slowly in this medium. It is interesting to note that the complex $[(C_6H_5COO)_2Ag]^-I^+$ reacts slowly with excess butadiene to form a syrup which upon hydrolysis gives an 80 per cent yield of 1,2-dihydroxy-3-butene (26). This reaction also proceeds by the mechanism given above, since the intermediate,



$$CH_2{=}CH\underset{\substack{|\\ O\\ |\\ C_6H_5C{=}O}}{C}HCH_2I$$

can be isolated.

It has been previously indicated that the halogen in the complex intermediate has positive character. Further proof of this is shown by the fact that it reacts with the alkynes, acetylene and phenylacetylene, by substitution rather than addition to give the corresponding haloalkynes (24).

B. The action of salts of dicarboxylic acids

Birnbaum and Gaier (3) first studied the reaction between equivalent quantities of the silver salts of some dicarboxylic acids and iodine. When silver oxalate is heated with iodine the only products obtained are carbon dioxide and silver iodide; salts of malonic, succinic, malic, tartaric, fumaric, and maleic acids yield the corresponding acids or anhydrides, carbon monoxide, carbon dioxide, and silver iodide. The reactions proceed vigorously, in the absence of air, below 200°C. Wieland and Fischer (32) reported results on the same compounds which differ slightly from those given above.

Wieland and Fischer and other investigators (10, 32, 33, 34) demonstrated that with silver salts of dicarboxylic acids capable of forming lactones, the chief products are the lactone and the acid or acid anhydride. For example, silver glutarate and iodine react to yield, besides silver iodide, γ-butyrolactone and glutaric anhydride.

C. The action of salts of α-hydroxy acids

As a general rule, salts of α-hydroxy acids react with equivalent quantities of iodine to give, as organic products, the free acid and the next lower aldehyde (13, 32). In this connection, it is interesting that only in the case of silver benzilate was an intermediate complex of the type discovered by Simonini (30) isolated (32). Apparently, the temperatures for formation and decomposition of most of the complexes of this type are very close together.

II. REACTIONS INVOLVING ONE EQUIVALENT OF SILVER SALT WITH TWO OF HALOGEN (SEE TABLE 2)

A. The degradation of the reaction product

When 1 equivalent of a silver salt of a carboxylic acid is treated with 1 mole of halogen in the cold, there is apparently formed a product of the composition RCOOX(1, 5), the nature of which will be discussed later in this report. Several investigators have studied the decomposition of this substance, RCOOX, where the halogen is either chlorine or bromine. In every case, there is formed by decarboxylation, the corresponding halo compound, RX, in excellent yield.

$$\mathrm{RCOOAg} + \mathrm{X_2} \longrightarrow \mathrm{RCOOX} + \mathrm{AgX}$$

$$\mathrm{RCOOX} \xrightarrow{\text{heat}} \mathrm{RX} + \mathrm{CO_2}$$

The salts of substituted acids react in a similar manner, giving always compounds poorer by one carbon atom.

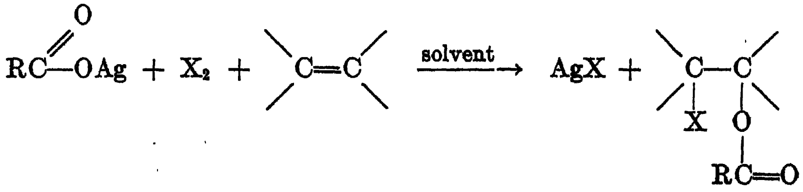
With a few exceptions, the compounds investigated have been salts of fatty acids or substituted fatty acids. Bockemüller and Hoffmann (5) have shown that bromine reacts with silver benzoate in boiling carbon tetrachloride to give an 80 per cent yield of bromobenzene. On the other hand, Lüttringhaus and Schaede (17) claim that with silver salts of aromatic carboxylic acids the degradation is unsatisfactory. The types of reactions carried out have been summarized (16). In each case the salts of the acid were used.

1. Alkyl halides are obtained from fatty acids.
2. α-Hydroxy acids give hydrogen halide compounds of the aldehydes.
3. From hydroxy acids in which the hydroxyl group is not in the alpha position, the hydroxyalkyl halides are formed.

4. Polyhydroxy acids with one hydroxyl group in the alpha position yield the hydrogen halide compounds of the hydroxyaldehydes.
5. 1,1-Halogenated hydrocarbons are formed from α-halo fatty acids.
6. From *n*-fold halogenated fatty acids the ($n + 1$)-fold halogenated hydrocarbons are obtained.
7. Dicarboxylic acids give rise to disubstituted hydrocarbon halides.
8. From ester acids the halogenated fatty acid esters are obtained. In this regard, it is noteworthy that Hunsdiecker, Erlbach, and Vogt (14) have utilized the ω-halo fatty acids, obtained by the degradation of the silver salts of dicarboxylic acid monoesters, to synthesize large-membered lactone rings and ring ketones.
9. α-Keto acids yield the acid halides poorer by one carbon atom.
10. Cycloalkane-monocarboxylic and -polycarboxylic acids yield the corresponding cycloalkyl halides.
11. From α-amino acids the alkylidenimine hydrohalides are obtained.
12. From other amino acids the haloalkylamines are formed.

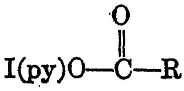
B. The reaction of the product with some nucleophilic substances

The product of reaction of equimolecular quantities of halogen and silver salt adds to the double bond of olefinic substances to give the corresponding halo esters (1, 5, 31). The general reaction may be expressed by the equation:



According to Birckenbach and coworkers (1), the immediate product of the reaction may be considered to be RCOOX, a mixed-halogen-like material. These investigators showed that the filtrate from the reaction at −80°C. of equimolecular quantities of silver acetate and iodine adds to cyclohexene at room temperature to give the acetate of 2-iodo-1-cyclohexanol. This is undisputed evidence for the presence of CH_3COOI in solution.

That iodine is undoubtedly the positive part of substances of the type RCOOI is shown by the work of Carlsohn (6), who was able to stabilize the positive iodine by coördination with pyridine, to form compounds of the type



$$\mathrm{I(py)O{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}R}$$

These compounds liberate iodine from potassium iodide, and in the presence of water undergo an autoöxidation–reduction, forming iodine and pyridinium iodate. Moreover, they immediately attack an acidic solution of phenol

TABLE 2

Summary of specific reactions for equimolecular quantities of silver salt and halogen

SILVER SALT OF ACID	HALOGEN	REMARKS	ORGANIC PRODUCTS	REFERENCES	YIELD
Acetic acid	Bromine	Dry state	Methyl bromide	(16)	80 per cent
Butyric acid	Bromine	Nitrobenzene as solvent	Propyl bromide and butyric acid	(5)	61 per cent of propyl bromide
Lauric acid	Bromine	Carbon tetrachloride as solvent	Undecyl bromide	(16)	75–80 per cent
Palmitic acid	Bromine	Dry state	Pentadecyl bromide	(16)	75–80 per cent
Stearic acid	Chlorine	Dry state; *ca.* 100°C.	Heptadecyl chloride	(16)	Variable
Stearic acid	Bromine	Carbon tetrachloride as solvent	Heptadecyl bromide and stearic acid	(5)	37 per cent
9,10-Dichloroöctadecanoic acid	Bromine	Carbon tetrachloride as solvent	1-Bromo-8,9-dichloroheptadecane	(16)	76 per cent
α-Bromostearic acid	Bromine	Carbon tetrachloride as solvent	1,1-Dibromoheptadecane	(16)	70 per cent
α-Hydroxypalmitic acid	Bromine	Dry state	The hydrobromide of the pentadecyl aldehyde	(16)	
Mandelic acid	Bromine	Ether as solvent	The hydrobromide of benzaldehyde	(16)	Variable
Pyruvic acid	Bromine	Carbon tetrachloride as solvent	Acetyl bromide	(16)	
Benzoic acid	Bromine	Boiling carbon tetrachloride as solvent	Bromobenzene	(5)	80 per cent
Phenylpropiolic acid	Iodine	Boiling benzene as solvent	1-Iodo-2-phenylacetylene	(32)	Good
Adipic acid	Bromine	Boiling carbon tetrachloride as solvent	1,4-Dibromobutane	(5)	10 per cent
Sebacic acid	Bromine	Carbon tetrachloride as solvent	1,8-Dibromoöctane	(16)	62 per cent
Adamantanedicarboxylic acid	Bromine	Carbon tetrachloride; *ca.* 70°C.	1,3-Dibromoadamantane	(21)	
Monomethyl ester of 1,6-hexanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 5-bromopentanoic acid	(15)	68 per cent
Monomethyl ester of 1,8-octanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 7-bromoheptanoic acid	(15)	70 per cent

Monomethyl ester of 1,9-nonanedioic acid	Bromine	Carbon tetrachloride solvent	Methyl ester of 8-bromoöctanoic acid	(15)	70 per cent
Monomethyl ester of 1,10-decanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 9-bromononanoic acid	(15)	75 per cent
Monomethyl ester of 1,11-undecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 10-bromodecanoic acid	(15)	71 per cent
Monomethyl ester of 1,13-tridecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 12-bromododecanoic acid	(15)	78 per cent
Monomethyl ester of 1,14-tetradecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 13-bromotridecanoic acid	(15)	71 per cent
Monomethyl ester of 1,15-pentadecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 14-bromotetradecanoic acid	(15)	73 per cent
Monomethyl ester of 1,16-hexadecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 15-bromopentadecanoic acid	(15)	70 per cent
Monomethyl ester of 1,17-heptadecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 16-bromohexadecanoic acid	(15)	70 per cent
Monomethyl ester of 1,18-octadecanedioic acid	Bromine	Carbon tetrachloride as solvent	Methyl ester of 17-bromoheptadecanoic acid	(15)	75 per cent
Acetic acid	Iodine	Ether solution of cyclohexene; room temperature	Acetate of 2-iodo-1-cyclohexanol	(1)	80 per cent
Acetic acid	Chlorine	Carbon tetrachloride solution of allyl chloride	Acetate of 2,3-dichloro-1-propanol	(5)	10 per cent
Propionic acid	Bromine	Chloroform solution of cyclohexene; −25° to −30°C.	Propionate of 2-bromo-1-cyclohexanol	(31)	48 per cent
Benzoic acid	Chlorine	Carbon tetrachloride solution of cyclohexene; −10°C.	Benzoate of 2-chloro-1-cyclohexanol	(31)	15 per cent
Benzoic acid	Bromine	Same as above	Benzoate of 2-bromo-1-cyclohexanol	(31)	42 per cent
Benzoic acid	Iodine	Ether solution of cyclohexene	Benzoate of 2-iodo-1-cyclohexanol	(1)	60 per cent
Phthalic acid	Iodine	Same as above	Di(1-iodocyclohexyl) phthalate	(1)	60 per cent
Benzoic acid	Iodine	Inert solvent containing pyridine	Monopyridine iodine (I) benzoate	(6)	Good
β-Naphthoic acid	Iodine	Same as above	Monopyridine iodine (I) β-naphthoate	(6)	Good
Benzilic acid	Iodine	Same as above	Monopyridine iodine (I) benzilate	(6)	Good

to give iodophenol, a reaction which Cofman (7) has shown can only be carried out by positive univalent iodine.[1]

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[1] When iodine or bromine is attached to a group which is more electronegative than a carboxylic acid residue, it is sufficiently positive (electrophilic) to iodinate or brominate benzene directly. Birckenbach and Goubeau (Ber. **65**, 395 (1932)) have demonstrated that iodine or bromine reacts with silver perchlorate in benzene to give a good yield of iodo- or bromo-benzene. They postulate the existence of $IClO_4$ as an intermediate in the first reaction.

THE SYSTEM $CaO–SiO_2–H_2O$ AND THE HYDRATION OF THE CALCIUM SILICATES

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I. INTRODUCTION

Many investigations of the system $CaO–SiO_2–H_2O$ have been made at ordinary temperature (15–30°C.), particularly during the last twenty years. However, the nature and disconcerting variety of the results make interpretation uncertain.

The difficulties arise mainly because the hydrous lime–silica products are gelatinous colloids. With change in the concentration of CaO in solution, the ratio CaO/SiO_2 for the solid phase continually alters instead of maintaining fixed values, as is usual for crystals of pure compounds of microscopic size. Accordingly, pronounced adsorption or some form of solid solution has generally been assumed. Water content, though not studied so extensively, is indicated to vary continuously with change in vapor pressure.

An abrupt change in the ratio of CaO to SiO_2 for the solid, with little change in the CaO in solution, is found at fairly low CaO concentration (0.04–0.15 g. per liter). Since the molar CaO/SiO_2 ratio for the solid increases in this region until it is 0.85–1.1, formation of $CaO \cdot SiO_2 \cdot aq$ is commonly assumed. The further, more gradual increase, that suggests adsorption or solid solution, raises the value to 1.4–2 at the CaO concentration corresponding to saturation with

$Ca(OH)_2$ (1.26–1.13 g. CaO per liter at 15–30°C.). The nature of this increase and the variation in the final value have both contributed to a conflict of opinion on the final state of actual chemical combination. Even the initial, more rapid rise in CaO/SiO_2 is variously interpreted. Several investigators believe that at low CaO concentrations $CaO \cdot 2SiO_2 \cdot aq$ is formed. Other compounds besides $CaO \cdot SiO_2 \cdot aq$ and $CaO \cdot 2SiO_2 \cdot aq$ have also been suggested. Though a number of investigators have drawn rather definite conclusions, the conflict in evidence and in opinion extends even to the most recent papers.

Many of the experimental studies have been made with the object of explaining the action of water on the tricalcium silicate, $3CaO \cdot SiO_2$, and β-dicalcium silicate, β-$2CaO \cdot SiO_2$, of portland cement. However, very dilute suspensions have generally been used, with frequent or continuous agitation, whereas cement itself is used in commercial practice as a plastic paste, without agitation, except initially. Such studies as have been made on pastes of the pure calcium silicates indicate, in general, a somewhat higher-limed hydrolysis product (CaO/SiO_2 about 2, or more) than has been obtained at the corresponding $Ca(OH)_2$ concentration (saturation) in most of the work with dilute suspensions. Thus, explanation of such differences constitutes an additional problem.

Owing to the confusing variety of results and opinions shown by the literature on the system $CaO–SiO_2–H_2O$, a critical, detailed review has been prepared. The aim has been to present an integrated account without, however, omitting details and discussions that show the actual differences in the data and opinions of the investigators. Though some of the details will be of interest mainly to the investigator contemplating further work in this field, the attempt was made to cover such matters rather thoroughly. There is much need for further, penetrating investigations made with full knowledge of what has already been found, a knowledge which, because of the state of the subject, is not attainable simply by study of the latest experimental papers. This review deals primarily with equilibria at ordinary temperatures. Of previous reviews, the more comprehensive and critical are those of Kühl and Mann in 1934 (97) and Bessey in 1938 (11).

II. A LITTLE EARLY HISTORY

For historical background it suffices to start with Le Chatelier's (117) classic dissertation on hydraulic mortars, published in 1887. Le Chatelier observed, as had others, that limewater and silica sol give a voluminous gelatinous precipitate. The molar CaO/SiO_2 ratio for the solid as first precipitated was 1.3, but the value was found to increase to 1.7 after six months of contact between precipitate and saturated limewater. Washing with water gradually removed CaO, until the ratio was about 1. Upon further washing, the successive portions of wash water came to about the same final concentration. Continued washing finally left only a residue of hydrous silica. Analysis of a precipitate that had been washed until the contacting solution remained only a little higher in CaO than the concentration (0.052 g./l.) at the apparent invariant point gave $1.06CaO:1SiO_2:2.47H_2O$. The method of drying was not stated.

Le Chatelier concluded that a compound, $CaO \cdot SiO_2 \cdot 2.5H_2O$, was formed which held additional lime by adsorption but remained up to lime saturation of the solution as the monocalcium compound. This opinion he reiterated in 1919 (119). In concluding that only the monocalcium compound was formed, he appears to have reasoned not only from the experiments described but also by analogy from the chemistry of the barium compounds. He showed that barium hydroxide and silica sol give a precipitate which appears amorphous at first but transforms to visible crystals of $BaO \cdot SiO_2 \cdot aq$ within a day. He tried in vain to precipitate a dibarium silicate.

Le Chatelier (117) also observed the formation of fine needle crystals in a hydrating cement. They were destroyed by ammoniacal salts, leaving gelatinous flocs which he believed to be hydrous silica. Hence, he concluded that the needles were composed of $CaO \cdot SiO_2 \cdot 2.5H_2O$, though he had obtained only gelatinous, amorphous-appearing solids in his precipitation experiments. From this occurrence of needles he drew support for his theory that cements harden by intergrowth and adhesion of elongated crystalline hydration products.

Michaelis, author of the opposing gel theory of cement hardening, published comprehensive addresses in 1906, 1907, and 1909 (133), in addition to numerous earlier papers. He precipitated silica sol with limewater and found that up to 1.5 moles of CaO could be brought into apparent combination with 1 mole of SiO_2. His work up to 1906 led him to a belief in hydrous compounds, similar to known hydrated magnesium silicates, with molar $CaO:SiO_2$ values of 3:2, 5:4, 4:3, and 1:1. By 1909 he had revised his ideas and recognized a continuous variation of the ratio of CaO to SiO_2 in the solid with change in the CaO concentration of the solution. He appears, however, to have retained a belief in the 1:1 compound; according to his computations, cements that hydrolyzed in large volumes of water appeared to give this compound. However, he had previously demonstrated to his satisfaction that continuous *leaching* eventually removed all the lime, leaving only hydrous silica.

Jordis and Kanter (77) in 1903 precipitated hydrated alkaline earth silicates at about 100°C., using silica sol or gel in saturated solutions of the respective hydroxides. In this way, crystalline $BaO \cdot SiO_2 \cdot H_2O$ and $SrO \cdot SiO_2 \cdot H_2O$ were obtained. The calcium compound was not distinctly crystalline and was more difficult to obtain in a pure state. However, analyses are reported that agree well with $CaO \cdot SiO_2 \cdot H_2O$. The authors concluded that a silicate containing more than 1 mole of alkaline earth to 1 mole of SiO_2 could not be formed in contact with water. Later (1906), Jordis (76) reported that the calcium silicate hydrate was prepared more satisfactorily by aid of calcium acetate. The product, $CaO \cdot SiO_2 \cdot 2H_2O$, was said to be amorphous but to be accurately the metasilicate. In an advance statement (1905) (75) the product was reported as $CaO \cdot SiO_2 \cdot H_2O$, and was called crystalline. Also of interest in this connection is a statement by Lafuma (105) that Becquerel obtained a crystalline calcium silicate, in an early experiment, by reaction of potassium silicate and calcium acetate "through the medium of parchment paper." Apparently, however, no analysis was reported.

Staining tests applied by Keisermann (1909–10) (81) and Blumenthal (1914) (16) to portland cement and calcium silicate "melts," hydrating on microscope slides, indicated that some fine needles and the gel that formed were composed of calcium silicate. The composition $CaO \cdot SiO_2 \cdot aq$ was inferred on the evidence of the earlier investigators, primarily Jordis and Kanter (77).

There were also other early studies, some of which will receive attention later. However, no data appear to have been reported prior to 1928 that give in detail the relations between solid composition and CaO in solution. Microscopical studies and attempts to identify the products of cement hydration constitute much of the early work. For the most part these attempts are now of little direct interest, since the methods of estimating free $Ca(OH)_2$ were generally very imperfect. Rebuffat (149) and others (118) concluded in favor of $2CaO \cdot SiO_2 \cdot aq$, but later the monocalcium silicate hydrate, $CaO \cdot SiO_2 \cdot aq$, appears to have received the widest acceptance as the most important siliceous product of cement hydration.[1] The experimental foundation for this belief was provided mainly by the works that have been discussed. Le Chatelier's formula, with the $2.5H_2O$, appears frequently in the early literature.

In a 1928 review on the setting and hardening of portland cement, Bogue (17) concluded that exact information on the major reactions of hydrating cement had not been advanced materially since the publication of Le Chatelier's dissertation.

III. THE LATER WORK: POSSIBLE CAUSES OF DIFFERENCES IN RESULTS

Since 1927, papers dealing systematically with the problems presented by the $CaO–SiO_2–H_2O$ system have appeared at frequent intervals. A few represent hydrothermal studies made at temperatures above 100°C. These are treated only incidentally in this review. For the most part they have no direct or certain significance with respect to the phenomena observed at ordinary temperatures (99, 175). Of the other studies, those that give apparent equilibria over a range in CaO concentration are represented in figure 1,[2] which shows the molar CaO/SiO_2 values[3] for the hydrous solids plotted against concentrations of CaO (mainly $Ca(OH)_2$) in solution. As previously mentioned, there are broad general similarities, but also pronounced individual differences.

Some of the differences in results are attributable, no doubt, to the differences in temperature, though the 40°C. and 20°C. curves (Nos. 11 and 12, figure 1) of Kühl and Mann (97) are more nearly the same than are some of the others.

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[1] See p. 13 of reference 93. At a symposium on the setting and hardening of cements and plasters, reported in the Journal of the Faraday Society (January, 1919), H. Le Chatelier, C. H. Desch, and A. A. Klein assumed $CaO \cdot SiO_2 \cdot aq$; G. A. Rankin tentatively assumed hydrous silica and adsorbed lime.

[2] Data obtained by Kryagova (90) have not been used in figure 1 because the units are thought to be incorrectly given by the author. For CaO/SiO_2 of about 1 in the solids the CaO in solution is reported as the exceptionally small quantity 0.005 g. per liter in tests with agitation, and as the exceptionally large quantity 0.6 g. per liter in tests made without agitation.

[3] In many cases these ratios were not obtained from direct analyses. Assumptions involved are considered in Section VI.

Experimental methods differed, of course, and evidently influenced the results, since there was commonly no complete check on the attainment of equilibrium. Some of the analytical procedures are also open to criticism.

Possibly of considerable importance, however, is the fact that, by reason of the colloidal states of the solid phases, complete thermodynamic equilibrium was actually not attained in any case. The solids may have remained significantly different in colloidal particle size in different investigations and even in different experiments within a single investigation. The resultant differences in adsorptive capacity may have contributed materially to the differences in curve slopes and to the range in the values of CaO/SiO_2 reached at lime saturation of the solution.

Also, differences in solubility caused by differences in particle size within the colloidal range could have affected the locations of the abrupt rises that are shown by the curves. Solubility differences could have contributed also to the fact that, in general, these rises are not truly vertical. If the common interpretation of them is correct, they represent what would normally be invariant points in the phase diagram, where two solid phases coexist with a solution of fixed composition, a composition determined by the solubilities of the two solids. Hence, if these solubilities change, the CaO concentration of the solution changes. Finally, it is at least conceivable that metastable compounds may have formed in some cases.

These possibilities—namely, that differences in colloidal state of subdivision may have contributed to differences in results—have received little mention in the literature on this system. However, Roller and Ervin (152) gave consideration to the possible effect of particle size on solubility and concluded from the general lack of evidence of aging that fairly definite solubilities might be maintained. Krasil'nikov and Kiselev (88) mention the degree of dispersion of the initial precipitate as having perhaps contributed to the different "invariant" concentrations found by different experimenters.

It is to be recognized that opinions differ as to the probable extent of the influence of fine particle size on solubility (66, 186) but solubility increases of at least 100 per cent appear to be distinct possibilities (66).

IV. RESULTS CONSIDERED IN RELATION TO INITIAL MATERIALS

The curves of figure 1 are arranged on the basis of a classification according to initial materials, which is shown in table 1. In the various cases, the silica was supplied as silica sol or gel; in solution as sodium silicate or as tetraethyl silicate; in calcium silicate "melts"; or in the practically pure anhydrous silicates, $3CaO \cdot SiO_2$, and β-$2CaO \cdot SiO_2$. That it is possible, starting with such a diversity of initial materials, to reach final states that are apparently not widely different is evident from the curves numbered 2, 8, 12, and 14.

The abrupt rise shown at the highest lime concentrations in a number of the curves of figure 1 was evidently caused, in general, simply by separation of $Ca(OH)_2$ as the solutions became saturated with this compound. (The CaO concentrations corresponding to saturation with $Ca(OH)_2$ at various temperatures have been indicated in figure 1 by use of I.C.T. data.) In certain cases the solutions became supersaturated with $Ca(OH)_2$; except for its lower terminus, Hedin's (67) curve, No. 10, is entirely for supersaturated solutions. Bessey (11) obtained the final rise in his curve, No. 8, at a CaO concentration so much lower than the 1.24 g. CaO per liter for a pure saturated solution of $Ca(OH)_2$ at

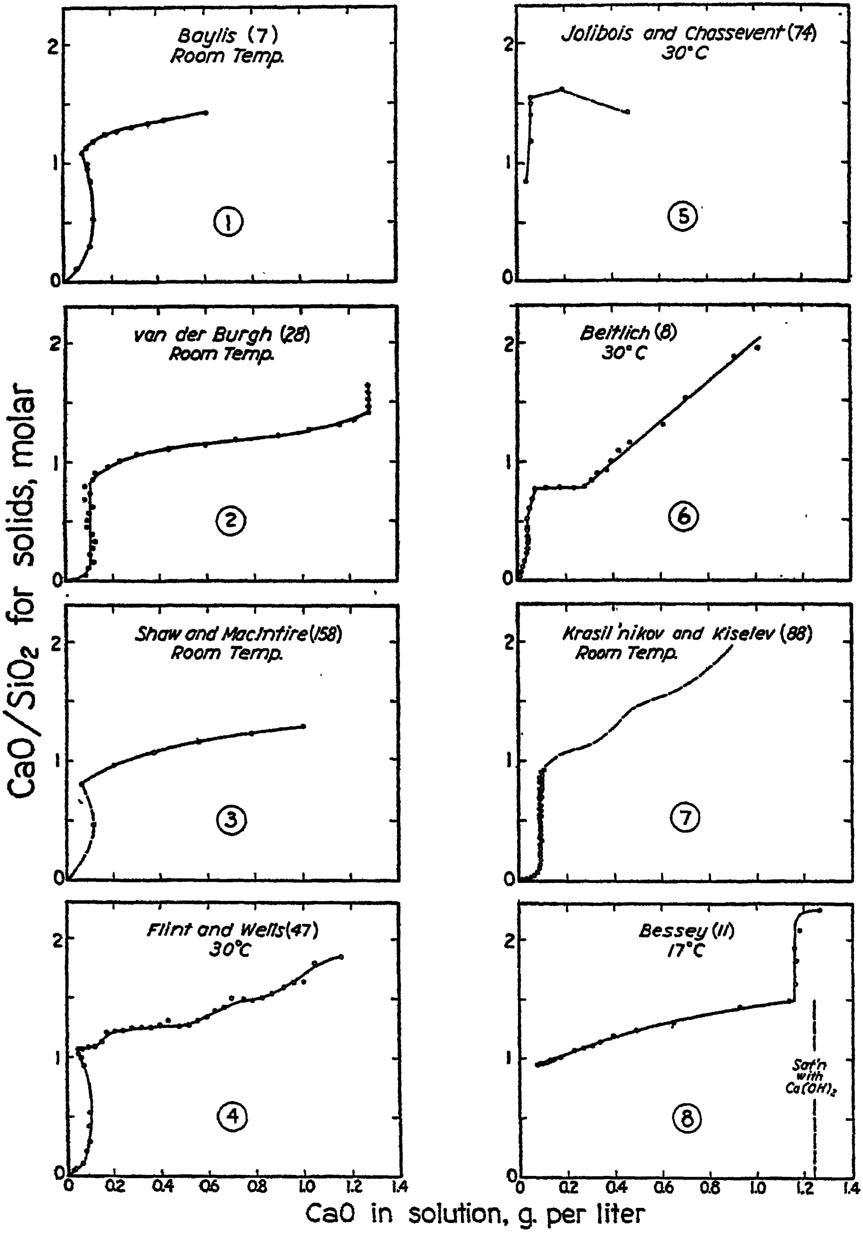

FIG. 1. The system $CaO-SiO_2-H_2O$; molar CaO/SiO_2 for solids *versus* CaO in solution

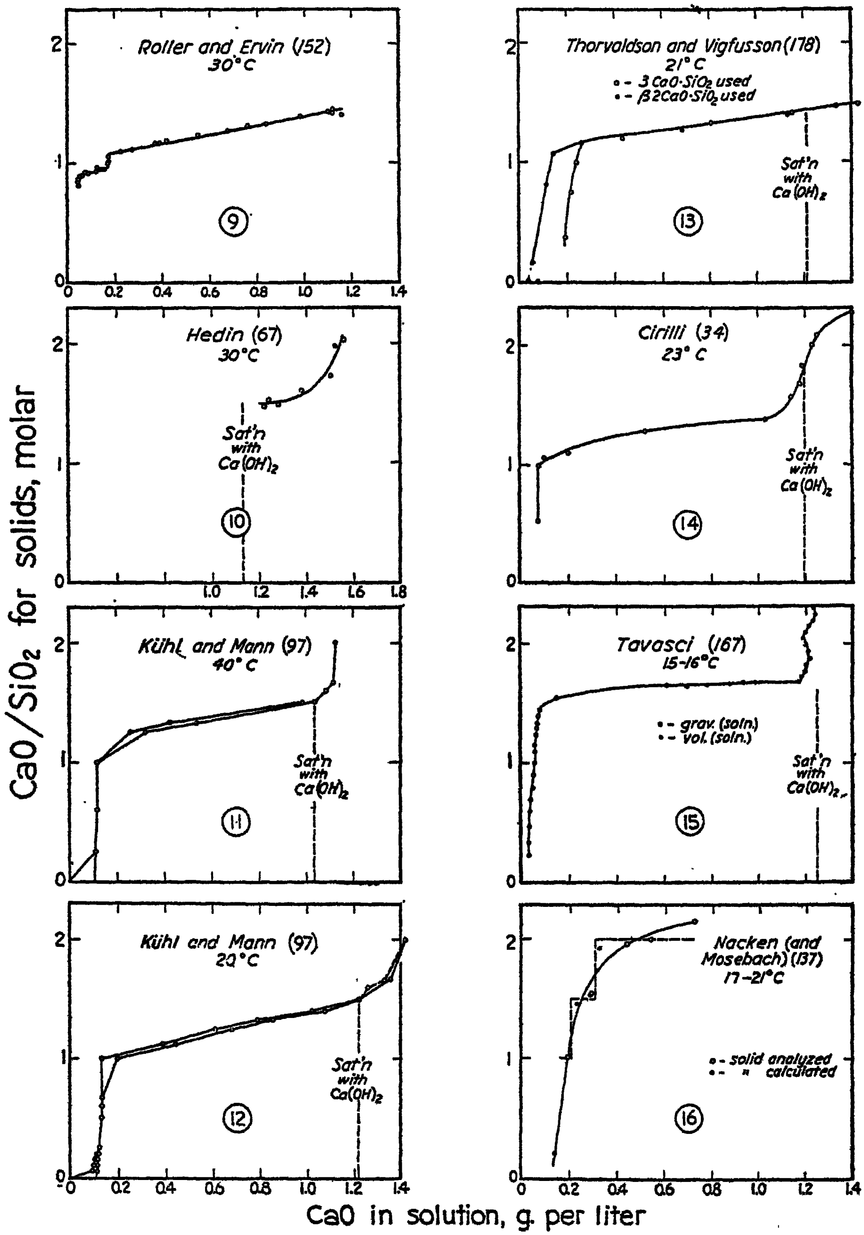

FIG. 1.—Continued

17°C. that he assumed separation of $2CaO \cdot SiO_2 \cdot aq$. Cirilli (35), however, has questioned the validity of this assumption. His own curve, No. 14, also begins to rise sharply before the CaO concentration is as high as for a pure saturated $Ca(OH)_2$ solution, but an x-ray spectrum showed $Ca(OH)_2$ in the solids (for $CaO/SiO_2 = 1.67$). Bessey obtained his data by adding water to a precipitate

TABLE 1

Initial materials used in the investigations represented in figure 1

IMMEDIATE SOURCE OF SILICA	INITIAL MATERIALS	CURVE NO.	TOTAL SiO_2	INVESTIGATORS
			grams per liter	
Silica sol or gel (including dried preparations)	Hydrous silica gel placed in water, and solid CaO added	1		Baylis (7)
		2	10	van der Burgh (28)
	Silica sol and limewater (for curve 3 a “solution-suspension”[2] of $Ca(OH)_2$ was used)	3	1	Shaw and MacIntire (158)
		4	0.45 or less	Flint and Wells (47)
		5	0.28	Jolibois and Chassevent (74)
	“Silicic acid”[2] (a C.P. reagent containing 20% H_2O) and limewater. Solid CaO was added in some cases.	6	0.6	Beitlich (8)
	Silica gel heated at 260°C., and later at 170°C.; retained 4.2% H_2O; used with limewater	7		Krasil’nikov and Kiselev (88)
Silicate solution	Sodium silicate and $Ca(NO_3)_2$; initial precipitate was $2.33CaO:1SiO_2:3.3H_2O$ and contained free $Ca(OH)_2$; dilute lime solution was added, and repeatedly diluted	8		Bessey (11)
	Sodium metasilicate and limewater; precipitate was placed in $Ca(OH)_2$ solutions of varied concentrations	9	About 0.06	Roller and Ervin (152)
	Tetraethyl silicate in alcohol solution was added to limewater	10		Hedin (67)
Lime–silica “melts”[2]	Quickly cooled melts of lime and silica were placed in water or limewater	11 12	10 or less	Kühl and Mann (97)
Anhydrous calcium silicates	$2CaO \cdot SiO_2$ and water	13	0.02–2.8	Thorvaldson and Vigfusson (178)
	$3CaO \cdot SiO_2$ and water	13	0.01–1.0	
	For curves 13 and 16, varied initial proportions of solid and water were used to obtain each point. For curve 15, the solution was replaced by pure water at intervals. No details were given for curve 14.	14		Cirilli (34)
		15	3	Tavasci (167)
		16	0.05–0.9	Nacken (137)

analyzing $2.33CaO:1SiO_2:3.3H_2O$ in which a part of the CaO was present as free $Ca(OH)_2$. However, if free $Ca(OH)_2$ still remained in the solid that is represented along the vertical part of Bessey's curve, his microscope failed to reveal it.

It is of some interest that Beitlich (8) and Krasil'nikov and Kiselev (88), who used the least hydrous silica gels, obtained the curves (Nos. 6 and 7) that rise the highest in figure 1 (except for the pronounced increases in CaO/SiO_2 at apparent saturation of the solutions with $Ca(OH)_2$). However, in both investigations, the solids were in contact with the solutions for months, as shown by table 2. Possibly this long contact may have as much significance as the state of the initial material, since the gels may be capable of very slow adjustments. Flint and Wells (47), whose tests lasted for a month, also obtained high values of CaO/SiO_2 (curve 4). As reported in Section II, Le Chatelier obtained a value of 1.7 in a six-month test. Also, Klasse and Kühl (84), who kept precipitates in saturated $Ca(OH)_2$ for three weeks, obtained CaO/SiO_2 values of about 1.8.

Shaw and MacIntire (158), in preliminary work, experimented with dried gels. By shaking with an attrition agent, they apparently attained nearly complete reaction within a few days. From these tests, at low lime concentration, they concluded that within wide limits the degree of hydration of the gel does not influence its capacity for absorption of $Ca(OH)_2$. Jordis and Kanter (77) found difficulty in obtaining definite alkaline earth silicates, whether of calcium, barium, or strontium, when their hydrous silica had been dried below 23 per cent water.

Curve 5, which Jolibois and Chassevent (74) obtained by mixing silica sols and limewater, stands out from its neighbors by reason of the high values of CaO/SiO_2 that were obtained, in this case at low lime concentration. For CaO/SiO_2 values up to 1.5, the CaO concentration is practically invariant at a value which in Le Chatelier's experiments seemed to indicate coexistence of $CaO \cdot SiO_2 \cdot aq$ and a lower-limed solid, perhaps $SiO_2 \cdot aq$. It appears that a particularly high degree of adsorption may be responsible for the unusual shape of curve 5. The tests were of six months' duration. As initially precipitated, all the solids, except the one for the first point, had CaO/SiO_2 values close to unity.

Curve 4, obtained by Flint and Wells (47), who also used silica sol and limewater, has a ragged-appearing upper section. However, since each point was obtained in a separate test conducted without agitation, this variability may well have no fundamental significance. Flint and Wells made other tests in which they shook anhydrous calcium silicates with water, filtered off the resultant supersaturated solutions, and then allowed the filtrates to stand, so that they precipitated. The precipitated solids had considerably higher values of CaO/SiO_2 than the solids of curve 4 at the higher CaO concentrations.

One peculiarity evident only in the curves for the silica sol and gel groups of table 1 is the bow-shaped initial rise whereby the curve turns back on itself. This odd effect will be discussed later.

Curve 9 was obtained by Roller and Ervin (152) by using hydrous lime–silica

TABLE 2

Experimental conditions in investigations represented in figure 1

CURVE NO.	INVESTIGATORS	PERIOD OF CONTACT IN OBTAINING INDIVIDUAL CURVE POINTS	AGITATION	EACH POINT A SEPARATE PREPARATION (?)	SPECIAL INDICATIONS OF EQUILIBRIUM; SPECIAL MEASURES TAKEN TO ESTABLISH EQUILIBRIUM
1....	Baylis (7)	2–3 days	Intermittent (one ⅝-in. steel ball used)	No	Auxiliary tests indicated reversibility. Author states equilibrium was nearly reached in 4 hr.
2....	van der Burgh (28)	24 hr.	Continuous	No	None, except duplicate tests which checked well
3....	Shaw and MacIntire (158)	5 days	Continuous	Yes	Change between 2 and 5 days was slight; there was some subsequent change, but it could be attributed to attack on glass vessels on basis of preliminary work
4....	Flint and Wells (47)	1 month	None	Yes	None
5....	Jolibois and Chassevent (74)	6 months	50 days continuous, then stood, except continuous agitation last 4 days	Yes	None
6....	Beitlich (8)	Up to several months	Continuous	Yes	Tests were continued until the electrical conductivity of the solution became constant
7....	Krasil'nikov and Kiselev (88)	Up to 17 months	Agitated, but no details given	Yes	Tests of varied duration showed attainment of constant concentration in the region of vertical rise in CaO/SiO_2 (figure 1); also, after additions of water or saturated limewater this concentration again established itself; the curve portion above the vertical rise was not well established, since in that region 17-month tests gave higher CaO/SiO_2 values than 10-month tests

TABLE 2—*Continued*

CURVE NO.	INVESTIGATORS	PERIOD OF CONTACT IN OBTAINING INDIVIDUAL CURVE POINTS	AGITATION	EACH POINT A SEPARATE PREPARATION (?)	SPECIAL INDICATIONS OF EQUILIBRIUM; SPECIAL MEASURES TAKEN TO ESTABLISH EQUILIBRIUM
8....	Bessey (11)	Not stated	Not stated	No	Author states "allowed to come to equilibrium"; no details
9....	Roller and Ervin (152)	0.5–8 days	Continuous	Yes (except one point)	Tests of varied length showed attainment of constant concentrations; points obtained in tests involving, respectively, absorption and release of CaO fall on the same curve; both kinds of test are represented in each curve section except the vertical rise at the lowest CaO concentration
10....	Hedin (67)	16 hr. (?)	Apparently none, except such as was caused by vacuum distillation	Yes	Vacuum distillation at 30°C. during addition of SiO_2 solution maintained practically constant $Ca(OH)_2$ concentration during reaction; in preliminary tests 16 hr. standing preceded analysis; whether final preparations stood is not stated
11....	Kühl and Mann (97)	13 days	Continuous	Yes	Curves were obtained from both the high- and the low-lime directions; see main text of this review
12....	Kühl and Mann	17 days	Continuous	Yes	As for curve 11
13....	Thorvaldson and Vigfusson (178)	About 10 days to 3 weeks	Continuous (metal shot used)	Yes	Separate tests of progressively increased length were run until the $Ca(OH)_2$ concentration remained constant
14....	Cirilli (34)	Not stated	Not stated	Apparently yes	No details

TABLE 2—*Concluded*

CURVE NO.	INVESTIGATORS	PERIOD OF CONTACT IN OBTAINING INDIVIDUAL CURVE POINTS	AGITATION	EACH POINT A SEPARATE PREPARATION (?)	SPECIAL INDICATIONS OF EQUILIBRIUM; SPECIAL MEASURES TAKEN TO ESTABLISH EQUILIBRIUM
15....	Tavasci (167)	Apparently several hours to several days	Continuous (glass shot used)	No	Test intervals were spaced according to the apparent needs as indicated by the continuity of results
16....	Nacken (137)	18–34 days	Continuous	Yes	For some compositions tests of varied duration were made

products precipitated from sodium metasilicate (and thoroughly washed before being suspended in $Ca(OH)_2$ solutions). It is unique in that it shows an additional, short, vertical rise at 0.17 g. CaO per liter. Such a "step" not only does not appear so clearly in any of the other curves, but it seems definitely not to be a possibility in some, because of close spacing of the experimental points. However, the rise is so small that the formation of a new compound, which it apparently indicates, might conceivably have occurred in other cases but have been obscured either by adsorption or by failure to establish equilibrium with sufficient accuracy. It is of interest that curve 9 was obtained at 30°C., midway between the two temperatures employed by Kühl and Mann (97). The latter investigators point out that their 40°C. curve, No. 11, rises rapidly after $CaO/SiO_2 = 1$ up to an apparent break at 1.25, though, as they say, the experimental points are too few for a close interpretation.

Two of the curves, Nos. 15 and 16, from the group of four that were obtained in investigations in which $3CaO \cdot SiO_2$ was used, show unusually high CaO/SiO_2 values at relatively low CaO concentrations. The nature of the initial material is such that results of this kind suggest incomplete equilibrium. This possibility, as it applies to curve 16, appears to be borne out in a subsequent report by Nacken (138), who (with Mosebach (137))[4] established curve 16.

Roller and Ervin have suggested that Beitlich may have obtained the high CaO/SiO_2 value that he did because of his use of solid lime, which may have become entangled in the gel. However, Baylis (7) and van der Burgh (28) also used solid lime, apparently with no ill effects.

Curves 13 were obtained by Thorvaldson and Vigfusson (178), by starting with anhydrous $3CaO \cdot SiO_2$ and $2CaO \cdot SiO_2$. Keevil and Thorvaldson (80) state that they obtained similar results by using crystalline hydrates of these same compounds, prepared hydrothermally. Bessey (11) employed crystalline hydrates in somewhat similar tests but found them too inactive for his purpose.

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[4] The writer has not had access to the paper by Nacken and Mosebach, but according to Building Science Abstracts 8, No. 1612 (1935) it covers the same investigation as Nacken's paper and in rather more detail.

In contrast, Keevil and Thorvaldson reported that the crystalline hydrates were more reactive than the corresponding anhydrous compounds. Bessey used, for his tests, water that already contained 0.4 g. CaO per liter, a point of difference cited by Thorvaldson in later discussion of Bessey's paper (173).

V. PROBLEM OF ATTAINMENT OF COLLOIDAL EQUILIBRIUM

With respect to most of the investigations represented in figure 1, some uncertainty exists as to whether equilibrium was actually reached, even on the colloidal level. Only in the investigations of Kühl and Mann (curves 11 and 12) and of Roller and Ervin (curve 9) was the curve obtained from both the high- and the low-lime directions over most of its course. However, Krasil'nikov and Kiselev established the vertical part of their curve (No. 7) in this way, making, in all, forty-two tests in this region.[5] Some of the other investigators took special pains to continue their tests until the concentration of the solution had been indicated to remain constant. The measures taken by the various investigators are summarized in table 2.

Duration of test varied widely, as shown. Roller and Ervin appear to have established that a few hours to a few days were sufficient for the adjustment of their lime–silica precipitates, whereas Krasil'nikov and Kiselev found ten months to be too short a time at the higher lime concentrations in their experiments with activated silica gel containing only 4.2 per cent of water.

Roller and Ervin report that equilibrium appeared to be attained more rapidly when the solids had to lose CaO than when they had to gain it. Tavasci (168) held the same view. Kühl and Mann found, at 20°C., a very pronounced difference of this kind at the point where CaO/SiO_2 for the solid was unity. Roller and Ervin found the time requirement to be relatively great at an apparent invariant point where two solid phases should be present.

Those investigators who sought to obtain equilibrium without or with only occasional agitation of their preparations were evidently influenced by the large volume occupied by the hydrous solid. Le Chatelier found that 1 g. of the precipitate occupied 2 liters. Baylis (curve 1) reported that his precipitates occupied half the total sample volume and that he therefore considered intermittent agitation to be adequate. Shaw and MacIntire (curve 3) observed that the settled volume increased regularly with increase in the CaO used; about 0.9 g. of SiO_2 added as sol to limewater (900 ml. total volume) gave sediments of 50 to 300 ml. after 16 hr. of settling.

The fact that Thorvaldson and Vigfusson (curves 13) obtained, over most of the range in CaO concentrations, the same curve for β-$2CaO \cdot SiO_2$ as for $3CaO \cdot SiO_2$ was to be expected if colloidal equilibrium was actually attained in each case, and if the hydration products were of equivalent texture. The difference between the curves at low lime concentrations was not considered significant by the authors, who state that the analytical precision was poor in that region. They state also that the tests on β-$2CaO \cdot SiO_2$ were made in a steel tube with steel shot whereas, in the region concerned, a gold-lined tube and gold shot were used with $3CaO \cdot SiO_2$, to prevent extraneous reaction. However, the results obtained with the steel tube and shot agree best with ones obtained by other investigators. The significance of the portions of Thorvaldson and Vigfusson's curves that correspond to low CaO concentrations is considered further in Section XVI.

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[5] Curve 7 of figure 1 was reproduced from microfilm; the many points in the region of the vertical rise could be reproduced only in a general way.

The drop in CaO concentration with increase in CaO/SiO_2 in solids above 0.5, that gives the bow shape, mentioned earlier, to the initial rise in some of the curves, strongly suggests either lack of complete colloidal equilibrium, non-maintenance of a fixed particle size, or analytical inaccuracy. Otherwise, it appears that such a reversal in curve direction should not occur. Whatever compounds are formed evidently constitute a series of hydrolysis products for which the CaO concentration would normally be expected not to decrease as CaO/SiO_2 increases.

Earlier it was noted that, since curves 15 and 16 were obtained by use of $3CaO \cdot SiO_2$, their unusual heights at moderately low CaO concentrations suggest the possible non-attainment of complete hydrolytic equilibrium. Nacken, who reported the data of curve 16, subsequently found (138) that in tests continued for nine months (instead of about 20–30 days as for curve 16) the CaO/SiO_2 values were markedly lowered. This was in spite of the fact that, as observed microscopically, most of the particles had appeared to hydrate clear to their centers within 3 days. Nacken considered that attack of the basic solution on the glass vessels had some influence on the long-time results but he came nevertheless to the conclusion that the short-time results probably represent only temporary equilibria. In spite of this, the relationships that Nacken supposed on the basis of the earlier results have subsequently been given considerable attention by others (39, 155), perhaps justifiably with respect to cement technology, since products formed during the first 20–30 days with consumption of a large part of the available water may tend to persist, whether they are equilibrium products or not. Nacken believed in the formation of rather definite stoichiometric compounds indicated by the dashed line in the figure. Schläpfer (155) notes that Nacken was unable to obtain similar results when using $2CaO \cdot SiO_2$. He reports that he, himself, has found that $2CaO \cdot SiO_2$ shaken with limewater containing 0.6 g. CaO per liter actually removes lime from solution, evidently by adsorption, since the $2CaO \cdot SiO_2$ gives the appearance of hydrolyzing.

Bessey (11) questioned whether curve 15 obtained by Tavasci represents equilibrium, since the methods used did not suffice to demonstrate equilibrium. Tavasci (168) defended his methods.

The work of Kühl and Mann (curves 11 and 12) is of special interest because of a unique method of investigation whereby, in spite of the fact that the solid particles hydrated only at their surfaces, the equilibrium relations appear to have been determined rather well. The investigators prepared melts of varied $CaO:SiO_2$ ratio and then sought to establish in each case the $Ca(OH)_2$ solution that would remain unchanged in CaO concentration as equilibrium was established at the surface of the solid. It was reasoned that if the amount of lime in solution remained unchanged, the equilibrium hydrate must have the same CaO/SiO_2

value as the melt itself. In practice the investigators were satisfied if the lime concentration changed only slightly. Indeed, at an invariant point, rather larger changes appear to have been considered satisfactory provided the invariant concentration was reached. Complete curves up to lime saturation[6] were established, both by starting with solutions more concentrated than the final ones, and *vice versa*. The near coincidence of these curves was believed to define with sufficient accuracy the course of the true equilibrium curve.

The principle of the procedure appears to be sound as long as the silica going into solution can be considered negligible. At the higher CaO concentrations, down almost to the concentration at the invariant point where the nearly vertical curve section is obtained, this assumption of negligible SiO_2 solubility is apparently justified. At the invariant point, however, there is reason to believe that the SiO_2 concentration is not insignificant. The SiO_2 removed from the solid must, then, affect somewhat the accuracy of determination of the terminal values of CaO/SiO_2 at this concentration. The degree of change in concentration permitted in attaining the invariant concentration would also affect the terminal CaO/SiO_2 values. However, in neither case should there be any effect on the accuracy with which the CaO concentration is determined at the invariant point.

The fact that the solid particles, though evidently very fine, hydrated only in a thin surface layer in spite of about 2 weeks of shaking with solution (2.1 g. solid in 70 ml. of solution) is largely due to the fact that the CaO/SiO_2 ratio was 2 or less. The $3CaO \cdot SiO_2$ hydrates and hydrolyzes much more rapidly. Thorvaldson and Vigfusson (178) show rate-of-hydrolysis curves (see later) that illustrate this difference for the β-$2CaO \cdot SiO_2$ and $3CaO \cdot SiO_2$ preparations with which they worked. They also report that, when they did not use their metal shot, *apparent* equilibrium, as determined by pH measurement and by titration, was attained while many of the solid particles still had unhydrated cores.

Hedin's experimental procedure (table 2), involving very slow precipitation with maintenance of constant $Ca(OH)_2$ concentration, appears to be well conceived.

VI. ANALYTICAL DETERMINATIONS

Just how dependent the differences in the various results shown in figure 1 may be upon inadequacies in the analytical work cannot now be fully established. However, a brief account will be given showing what determinations were made and giving a few additional details, mainly such as seem to have special significance. Unfortunately, most of the authors give little information regarding their analytical work.

The CaO in solution was determined volumetrically (by use of permanganate) for curve 9 and gravimetrically for curves 4 and 16, and for much of curve 15. Titration with acid was used for curves 11 and 12, but was checked to some extent by gravimetric tests. For curves 1, 2, 3, 7, and 13, the CaO concentrations were determined solely by acid titration. The methods of determining CaO concentrations for curves 8 and 14 were not stated, and the exact method used for curve 10 is uncertain, though the lime was apparently first precipitated as oxalate. For curves 5 and 6 the determinations were based on the electrical conductivity of the solution. When this determination is made simply by comparison with the conductivities of pure $Ca(OH)_2$ solutions, as was true relative at least to curve 6, it is subject to error at low CaO concentrations where the SiO_2 in solution becomes significant (152). In such tests, as well as in acid titrations, error can result also if alkalies are dissolved from glass vessels.

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[6] Both the solid and the dashed line shown in plot 12 of figure 1 for solutions supersaturated with $Ca(OH)_2$ were obtained from the low-lime side. The dashed line resulted when "dusted" melts containing some γ-$2CaO \cdot SiO_2$ were used.

Tavasci (curve 15) found by checking against gravimetric determinations that acid titration based on phenolphthalein indicator gave CaO values that were too low when the CaO was 0.14 g. per liter or less. Titrations based on methyl orange were shown to be satisfactory. Results obtained with sodium silicate are similar (21, 183). Baylis (curve 1) also noted the difference between phenolphthalein and methyl orange end points, and used the latter in calculations. Shaw and MacIntire (curve 3) apparently used only the phenolphthalein indicator. Krasil'nikov and Kiselev (curve 7) used bromothymol blue after checking its suitability by potentiometric titration. The indicators used for curves 2 and 13 were not stated.

The CaO/SiO_2 values for the solids were commonly not determined directly, except sometimes for the final composition after a series of interrelated tests. Instead, they were generally computed from final solution concentrations and the amounts and compositions of initial solids and solutions. Curves for which the CaO/SiO_2 values of the solid products were not in general obtained by direct analysis are 1–6 inclusive, 11–13 inclusive, 15, and apparently 7 also. For the points on curves 9 and 10, and some of those (as marked) on 16, direct analyses of the solids were made. Whether each solid was analyzed for curves 8 and 14 was not reported.

For many of the curves, as represented in figure 1, the CaO/SiO_2 values were computed by assuming all the SiO_2 to be present in the solid. This is true of curves 2, 3, 6, 11, 12, 13, and part of 16 (all the calculated points), while the situation regarding curves 1, 7, 8, and 14 is not clear. In general, the practice of the investigators has been conformed to, but not all of them drew curves of this type. Thorvaldson and Vigfusson plotted the initial concentrations of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ against "mols CaO retained." Their curves show abrupt breaks, or inflections, where the retained CaO is about 1 mole per mole of the silicate. Hence, points of inflection were said to occur "where the composition of the residue corresponds to monocalcium silicate." Neglect of the SiO_2 in solution is justifiable at the higher lime concentrations, and in general when the proportion of total SiO_2 to volume of solution is sufficiently large, but may have important consequences if the total SiO_2 is relatively small and the CaO concentrations are low. As the CaO concentrations drop below 0.1 g. per liter, the SiO_2 concentration increases rapidly (152). Thorvaldson and Vigfusson (curves 13) had to use relatively small quantities of solids to obtain the lower lime concentrations, since their procedure was simply that of adding water to the anhydrous calcium silicates. The possible consequences of this are considered in Section XVI.

The silica concentrations were determined for the solutions represented in curves 1, 4, 5, 7, and 15. Baylis (curve 1) and Krasil'nikov and Kiselev (curve 7) questioned whether the silica which they found was necessarily all in true solution. Roller and Ervin did not determine SiO_2 for the solutions used in establishing curve 9, but they did make SiO_2 determinations on similar solutions, separately prepared. Shaw and MacIntire (curve 3) also made supplementary SiO_2 determinations. Values obtained by the various investigators will be considered in Section IX. The determinations made by Roller and Ervin and by Krasil'nikov and Kiselev were colorimetric. Kryagova (90) has also made colorimetric determinations and gives considerable detail on his method.

Various investigators report precautions taken against the action of atmospheric carbon dioxide, but few appear to have determined the extent of carbon dioxide contamination. Tavasci (curve 15) found at the end of his tests, which involved many withdrawals of solution and replacements by water, that the carbon dioxide in the solids actually amounted to 45 per cent of the remaining SiO_2; he applied graduated corrections to the data obtained earlier. Roller and Ervin analyzed and found that only a small correction for carbon dioxide was needed in establishing curve 9. Hedin (curve 10) examined his products under the microscope and could observe no crystalline calcium carbonate. Bessey (curve 8) reported that he corrected for carbon dioxide.

Owing to the change in composition of the solid with change in the amount of CaO in solution, it is apparent that the washing of a precipitate preparatory to analysis could introduce error. Among early investigators, Jordis and Kanter (77) took precautions against such error by washing first with alcohol diluted 1:2, then with absolute alcohol, and finally with ether. Baylis (curve 1) restricted his volume of wash water to an amount about 50 per cent in excess of the volume of the precipitate. Roller and Ervin (curve 9) washed precipitates obtained at the higher lime concentrations with aqueous acetone "of such concentration (50% or stronger) that a precipitate of calcium hydroxide just was avoided." In a second wash the concentration of acetone was raised. Pure acetone was eventually used, followed by pure ether. Nacken, and Hedin, who also analyzed a number of hydrous products, do not give explicit details.

Shaw and MacIntire obtained evidence that when they used steel shot in a glass vessel results obtained after 2 days were apparently affected by attack upon the glass. (No shot were used in obtaining their curve 3.) Tavasci (curve 15), who used glass shot, concluded from the good agreement between his gravimetric and acidimetric determinations of CaO that contamination by alkali from the glass was negligible. Krasil'nikov and Kiselev (curve 7) made blank tests using only $Ca(OH)_2$ solution in their reaction vessels, which were glass ampoules. After seven months the more concentrated solutions had diminished in CaO by 7 per cent, but no SiO_2 was found in solution. In the weaker solutions the loss of CaO was inappreciable. Kühl and Mann (curves 11 and 12) considered that by using only glass flasks that had long been used for similar work, or new ones that had been well steamed, they largely avoided further leaching of the glass. Checks which they obtained between the results of acid titrations and gravimetric determinations for CaO appeared to confirm this. Thorvaldson and Vigfusson (curves 13) and also Bessey (curve 8) used metal vessels. Roller and Ervin (curve 9) made their supplementary solubility determinations by use of a paraffined reaction flask. Thorvaldson (173) and Bessey (12) have both expressed the opinion that the use of any kind of glass vessels is undesirable.

VII. CHEMICAL COMBINATION OF HYDROUS SILICA AND LIME

In the past it has occasionally been assumed that hydrous silica does not, or at least may not, undergo at room temperature any actual chemical combination with lime (23, 38, 145, 180). In view of the uncertainties that persist regarding the $CaO-SiO_2-H_2O$ system, and the resemblances between the properties of the hydrous lime–silica products and what might be expected of a gelatinous adsorptive system, such views are understandable. However, though serious question yet remains as to the state of combination of some of the lime in the higher-limed products, the evidence for chemical combination of the silica with at least a part of the lime now seems practically conclusive. Considerable evidence exists, as is shown by the following account.

The observations of various authors have led them to ascribe to lime a peptizing, or dispersing, action on silica gel, which action is apparently similar to that produced by sodium hydroxide (30). Maffei and Battaglia (127) and Cirilli (32) observed that after the initial rate of removal of lime from solution by silica gel had fallen off, the rate again became greater, an effect which they attributed to dispersing action of the lime and alteration of the original gel structure. Van der Burgh (28) reported that in his (shaking) experiments, the solid phase which was at first in the form of fragments of gel changed into a very voluminous, flocky, snow-like mass. Mehrotra and Dhar (128) reported that hydrous silica gel was easily taken into colloidal solution by bases, including $Ca(OH)_2$. Flint and Wells (47) found that silica gel dissolves slowly in dilute $Ca(OH)_2$ solution at 30°C.; at 100°C. the action was much more rapid and it gave a clear solution, a result obtained also by Kryagova (90). Relative to chemical reaction, these various observations are suggestive only, but they show at least that the action of the lime is such as to destroy the original gel structure.

More direct evidence for chemical combination is provided by the following additional observations: Mehrotra and Dhar found in their tests at room temperature that gelatinous silica acts acidic, in that it "adsorbs" bases to a much greater extent than acids. Kolthoff and Stenger (86) found that $Ca(OH)_2$ was "adsorbed" much more strongly than alkali hydroxides, and hence appeared to form a relatively insoluble silicate. Maffei (126) analyzed data obtained after contact periods of systematically varied length, and showed that the constants of the Freundlich and Langmuir adsorption isotherms vary in a way that suggests that CaO is first adsorbed by silica gel and then reacts chemically. Indeed, Maffei's analysis indicates also that the products are subject to hydrolysis. Krasil'nikov and Kiselev (88) show graphically that curves obtained at early time intervals are like ordinary adsorption isotherms but that a progressive change occurs until the vertical section shown by curve 7 of figure 1 is obtained. This was found to be true whether the initial material was silica gel or naturally occurring tripoli or diatomite.

This vertical or roughly vertical curve section that appears at low CaO concentration in plots of CaO/SiO_2 in solids *versus* CaO in solution is considered by many to be rather good evidence of chemical reaction. A truly vertical rise does, under ordinary circumstances, indicate the coexistence of two solid phases in a three-component system; in the present case one of these phases would be a calcium silicate. The departures from a strictly vertical curve section can be accounted for by the difficulties of experimentation and the colloidal nature of the solids, as discussed earlier. Such aspects can account also for the fact that the SiO_2, as well as the CaO, is found to vary in concentration. It may also be that more than one phase transition is involved.

Klasse and Kühl (84) sought to produce calcium silicate hydrates by double decomposition of sodium silicates and calcium nitrate at 19°C. They believed that they were successful in their efforts to obtain $CaO \cdot SiO_2 \cdot aq$ by use of pure sodium metasilicate. The molar CaO/SiO_2 ratio for the product was always somewhat less than 1, but the difference appeared to be caused by hydrolysis, since it was reduced by using excess calcium nitrate. In this way a ratio of 0.96:1 was attained. Emley (44) precipitated what he believed to be hydrated monocalcium silicate, using calcium chloride and sodium metasilicate. The washed and dried product showed no free CaO or $Ca(OH)_2$ when tested by Emley's glycerol–alcohol solution method (see Section XIII).

Cirilli (32) allowed a freshly precipitated silica gel (about 80 per cent water) to interact with saturated limewater and obtained for the product a heat of solution of 362.6 cal. per gram of CaO, as compared with 557.9 cal. per gram of CaO for $Ca(OH)_2$. Thus, under the assumption that any unreacted silica would not be materially affected by the acid, the heat released in the formation of the lime–silica product was about 11 Cal. per mole of CaO. The value is large enough to suggest chemical reaction.

That silica gel in aqueous suspension can interact with CaO at room temperature until it becomes apparently completely acid soluble was observed at least as early as 1883, being reported in that year by Landrin (106). Lerch and Bogue (122) observed that a sample of hydrous silica that had stood in saturated $Ca(OH)_2$ solution for a year had changed till 97 per cent of the silica was acid soluble. The soluble part of the product had the composition $1.07CaO:1SiO_2$. Shaw and MacIntire (158) reported that "numerous qualitative tests with solid residues, in which the factors of CaO/SiO_2 ratio and time were sufficient to assure the formation of $CaSiO_3$ have shown that this product is quickly and completely dissolved by cold dilute (1+9) HCl." Such solutions no doubt soon become colloidal, but apparently silicic acid can form initially in a state of true solution (69).

That definite, negatively charged silicate ions exist in appreciable concentrations in weakly alkaline solutions at room temperature is indicated by the work of many investigators (153). Moreover, as the $Ca(OH)_2$ concentration of the solution is raised, precipitation occurs and the SiO_2 concentration falls to a very low value (47). Thus, precipitation of a calcium silicate is strongly indicated. It is significant in this connection that hydrated barium silicate precipitates at room temperature, as Le Chatelier (117) showed.

Keisermann (81) found that methylene blue stains only free silicic acid and not the silicates. It gave a negative test with the gel produced by action of water on the anhydrous calcium silicates, except when in acid solution, in which case the test was positive. Accordingly, a silicate was believed to have been demonstrated. Blumenthal (16) confirmed these observations. Later, Diehl made similar tests (for Nacken (137)) on hydrating tricalcium silicate and also obtained similar results. As previously noted, Keisermann and Blumenthal also observed needle formations which staining tests (especially those by Blumenthal) indicated were calcium silicate.

Richter and Radczewski (150) found that silica sol that had adsorbed lime showed initially no electron-diffraction pattern attributable to a crystalline product, but soon gave such a pattern under the continued action of the electron radiation. Eitel (40) concluded from this that the products of cement hydration pass only temporarily through the "colloid-amorphous" state and soon become crystalline though highly disperse, i.e., crystalline colloids.

Chassevent (31) examined by x-ray the products of the interaction of limewater and silica sol, at ordinary temperature, and obtained within about 2 hr. an x-ray spectrum which remained, within the limits of accuracy of the observations, unchanged for products of varied CaO/SiO_2 ratios (a characteristic confirmed by Cirilli (35)). Later, Mlle. Foret (49) prepared (hydrothermally, below 140°C.) crystalline $CaO \cdot SiO_2 \cdot aq$, the composition of which she verified by chemical analysis, and which gave an x-ray spectrum which she considered identical with that obtained by Chassevent. Strätling (165), Brandenberger (19), Forsén (53), and Cirilli (35) have obtained a similar x-ray spectrum for hydrous lime–silica products prepared at ordinary temperature. Except for Brandenberger, who did not express an opinion, these latter authors believed their compounds to be more highly limed than $CaO \cdot SiO_2 \cdot aq$. Whatever the actual compounds or solid solutions that may be formed, the x-ray tests alone appear to be strong evidence that the lime does combine chemically with the silica at ordinary temperature. However, to recapitulate, the evidence for compound formation may be summarized as follows:

Lime acts on silica gel with destruction of the gel's original structure. The lime taken up gives, at first, a characteristic adsorption curve, but this curve changes with time until it strongly suggests the existence of an invariant point denoting the presence of two solid phases. The attendant heat evolution is large enough to suggest chemical reaction. Also, the silica in solution decreases markedly with increase in lime concentration, consistent qualitatively with maintenance of a solubility product for calcium silicate. Moreover, the silica in the solid becomes completely soluble in a moderate volume of acid. Staining tests also indicate that the silica is combined, and one of the better tests for free CaO or $Ca(OH)_2$ has given a negative result when applied to a precipitate believed to contain as much lime as silica. Under electron radiation an electron-diffraction spectrum soon appears. Examination by x-ray gives a diffraction spectrum that is apparently the same as one produced by microscopically crystalline hydrated calcium silicate.

VIII. COMPOUNDS FORMED AT LOW LIME CONCENTRATIONS

The curves of figure 1 that begin at the origins of the plots show, in general, an initial short section where the increase in the CaO/SiO_2 value is less rapid than for the abrupt rise which follows. In this initial region the solid phase is commonly believed to be simply the hydrous silica, the lime being merely adsorbed (8, 47, 88, 171). However, the more nearly vertical curve section that follows suggests chemical reaction, as already noted.

Hydrous calcium silicates that have been assumed at low lime concentrations, tentatively or otherwise, by the authors whose data are shown, in part, by figure 1, are listed in table 3. In most cases, the constitutions of the compounds were suggested simply by the points at which breaks in the experimental curves were observed. However, a degree of judgment was usually exercised in arriving at simple combining proportions. In some cases the CaO/SiO_2 value for the assumed compound is less than for the experimental point, a difference which can be attributed to adsorption. In other cases the assumed ratio is greater than the experimental one. Unless the difference can be attributed to analytical inaccuracy, it would seem that failure to obtain complete equilibrium would be the most plausible explanation. Van der Burgh (28) suggested adsorption to account for the fact that his curve rounds off before CaO/SiO_2 reaches 1,[7] but he appears to have meant simply that adsorption proceeded as though the SiO_2 had completely reacted whereas some remained, evidently in poor contact with the solution.

TABLE 3

Compounds that may exist at low lime concentrations (investigations represented in figure 1)

CURVE NO.	INVESTIGATORS	COMPOUNDS ASSUMED OR SUGGESTED			
		$CaO \cdot 2SiO_2 \cdot aq$ (1:2)	$3CaO \cdot 4SiO_2 \cdot aq$ (3:4)	$4CaO \cdot 5SiO_2 \cdot aq$ (4:5)	$CaO \cdot SiO_2 \cdot aq$ (1:1)
1	Baylis (7)				x
2	van der Burgh (28)				x
3	Shaw and MacIntire (158)		x		x
4	Flint and Wells (47)	x			x
5	Jolibois and Chassevent (74)				x
6	Beitlich (8)	x		x	
7	Krasil'nikov and Kiselev (88)				x
8	Bessey (11)				x
9	Roller and Ervin (152)	x	x		x
11–12	Kühl and Mann (97)				x
13	Thorvaldson and Vigfusson (178)				x
14	Cirilli (34)				x
15	Tavasci (167)	x			x
16	Nacken (137)				x

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[7] Of somewhat similar nature is the suggestion by Kühl and Mann (100) that adsorption may account for the fact that they did not obtain a curve section strictly perpendicular to the CaO axis.

Flint and Wells (47), Roller and Ervin (152), and Cirilli (32) had other reasons for assuming the compound $CaO \cdot SiO_2 \cdot aq$. Flint and Wells determined both SiO_2 and CaO in solution and plotted one against the other, with the CaO/SiO_2 value for the solids also shown. A change in curve direction in a region where this CaO/SiO_2 value increased rapidly from 0.537 to 0.930 was assumed to denote formation of $CaO \cdot SiO_2 \cdot aq$.

Roller and Ervin tested their sloping curve sections for agreement with the Freundlich adsorption isotherm, by assuming adsorption on hydrated calcium silicates of various molar ratios. For the curve section at the higher CaO concentrations the agreement was best, and indeed very satisfactory, when they assumed adsorption on $CaO \cdot SiO_2 \cdot aq$. As a further test they estimated ion activities and showed that $CaO \cdot SiO_2 \cdot aq$ apparently maintains a constant activity solubility product within the same region.

Cirilli determined heats of solution after exposing hydrous silica for successively increased times to saturated $Ca(OH)_2$ solution. The values per gram of CaO remained constant until CaO/SiO_2 reached 1; then they increased.

Curves 3, 5, 15, and 16 of figure 1 that are due to Shaw and MacIntire (158), Jolibois and Chassevent (74), Tavasci (167), and Nacken (137) do not, themselves, show clearly why the authors believed in $CaO \cdot SiO_2 \cdot aq$. Shaw and MacIntire suggested that their curve might not have represented complete equilibrium. Jolibois and Chassevent were influenced by the fact that their precipitates had, initially, $CaO:SiO_2$ values close to 1:1; also, by the fact that the apparent invariant concentration, 0.053 g. CaO per liter, checked that given by Le Chatelier (117), whose work suggested the 1:1 compound. Tavasci was guided also by Le Chatelier's value of 0.052 g. CaO per liter; he found 0.053 g. CaO per liter when CaO/SiO_2 in solids was about 1. He also found that the CaO concentration was the same for the two CaO/SiO_2 values 1.0 and 1.1. Nacken was influenced by the belief that he was obtaining a succession of definite compounds (see dashed line in curve 16, figure 1).

As mentioned in Section II, Jordis and Kanter appear to have produced a hydrated monocalcium silicate, but it was obtained at about 100°C. Le Chatelier's evidence, previously presented, is obviously more suggestive than conclusive with respect to exact compound composition. Schott (157) experimented with portland cement, using a special technique, and also obtained a colloidal product whose analysis was closely that of $CaO \cdot SiO_2 \cdot aq$. Hedin (67), working with Forsén (52), found that $3CaO \cdot SiO_2$ shaken with a large volume of water gives a voluminous precipitate analyzing initially 1.05–1.13 CaO to 1 SiO_2.

It will be recognized that some of the evidence on which the compound $CaO \cdot SiO_2 \cdot aq$ has been assumed is very weak and much of the rest merely suggests a compound having a $CaO:SiO_2$ ratio of *about* 1:1. The method of Roller and Ervin (based on the solubility product) appears somewhat more decisive, but it is dependent on acid-dissociation constants that are not well established (see Section X). Cirilli's heat data seem significant, but their simplicity may be deceptive. That is, since the solution was kept saturated with $Ca(OH)_2$, the equilibrium product toward which the reaction tended was always one having a

$CaO:SiO_2$ ratio considerably greater than 1:1. It would seem that either adsorption or formation of a compound higher limed than 1:1 tended to occur. Why the curve should break precisely at the 1:1 point is thus left somewhat obscure. On the other hand, Cirilli later (34) reported heat data for hydrolysis products of $3CaO \cdot SiO_2$ which show a break at the same 1:1 composition.

The x-ray studies mentioned in Section VII appear to be as good evidence as any for the formation of $CaO \cdot SiO_2 \cdot aq$ at ordinary temperature. Mlle. Foret's crystals, analyzed as $CaO \cdot SiO_2 \cdot aq$, gave the same spectrum as Chassevent's room-temperature preparations. Moreover, Mlle. Foret later (50) tried a CaO to SiO_2 ratio of 2 and found that the hydrothermal products obtained at 130°C. or less were still $CaO \cdot SiO_2 \cdot aq$, plus free CaO, and gave the same spectrum for the hydrated calcium silicate as before. However, as mentioned earlier, certain other investigators who obtained practically the same spectrum believed that it represented a compound higher limed than $CaO \cdot SiO_2 \cdot aq$.

Kühl and Mann (97), after reviewing the literature up to 1934, concluded that the formation of $CaO \cdot SiO_2 \cdot aq$ (presumably at room temperature) could be considered as established. In 1938 Thorvaldson (173) did not regard the matter as fully settled but Bessey (12) believed that there was little room for doubt.

Several experimenters have concluded that their data show that compounds lower in lime than $CaO \cdot SiO_2 \cdot aq$ are not formed at room temperature. Authors who make a definite statement to this effect are Kühl and Mann (102), Cirilli (32), and Krasil'nikov and Kiselev (88). Cirilli based his opinion on his evidence, already mentioned, that when silica gel is exposed to a solution kept saturated with $Ca(OH)_2$ the heat released per gram of CaO taken up remains constant until the ratio CaO/SiO_2 reaches 1. However, the fact that the solution was kept saturated with $Ca(OH)_2$ would seem to destroy the value of the evidence with respect to equilibrium at low lime concentration. Cirilli's (34) later heats of solution for hydrolysis products of $3CaO \cdot SiO_2$ cover only one composition that has a CaO/SiO_2 value below 1. The other authors relied on the fact that their data indicated that solutions in equilibrium with solids for which the CaO/SiO_2 value varied from about 0 to nearly 1 were almost invariant in CaO concentration. Van der Burgh (28), who obtained similar results, is not as explicit but appears to have been confident that $CaO \cdot SiO_2 \cdot aq$ is the first compound formed. Bessey (11) pointed out that the experiments of Lerch and Bogue (122), in which leaching of $3CaO \cdot SiO_2$ left a residue of hydrous silica, indicate that $CaO \cdot SiO_2 \cdot aq$ must dissolve incongruently, whereas Flint and Wells' (47) interpretation of their curve requires congruent solution. This discrepancy had been reported previously by Lea and Desch (114; see also 137). Bessey concluded from it that all of Flint and Wells' solubility curve up to a point at which CaO/SiO_2 for the solution exceeds 1 represents metastable conditions. Hence, he concluded that under equilibrium conditions $CaO \cdot SiO_2 \cdot aq$ is the lowest-limed silicate hydrate; this, however, is not a necessary conclusion from the Lerch and Bogue experiments.

The experimenters who expressed a belief in compounds lower limed than $CaO \cdot SiO_2 \cdot aq$ generally relied mainly on breaks in their curves. Roller and

Ervin (152) had no direct experimental data of their own for $CaO \cdot 2SiO_2 \cdot aq$, but considered that the solid next below their assumed compound $3CaO \cdot 4SiO_2 \cdot aq$ could not be silica. They believed that the solubility of gelatinous silica is greater than the concentration which they found at the invariant point in question. These authors found that their Freundlich isotherm test was indecisive with respect to $3CaO \cdot 4SiO_2 \cdot aq$. The decision in favor of this compound was made on the basis of the solubility product test. Assumptions were involved which will be indicated in Section X. Flint and Wells' solubility data (47) indicated an abrupt decrease in SiO_2 concentration beginning at the point where CaO/SiO_2 for the solid first reached 0.1. Though the solid did not attain a CaO/SiO_2 value of 0.5 until the SiO_2 concentration had dropped very materially, it was thought that equilibrium was probably not fully attained and that the break in the curve represented formation of $CaO \cdot 2SiO_2 \cdot aq$. The authors were apparently influenced also by consideration of the ions with one charge per silicon atom which their mathematical analysis suggested were present in solution (see later). There is of course other evidence also that such ions exist in solutions of moderate alkalinity (73, 181, 182).

To summarize, apparently one or more hydrated calcium silicates are formed at low lime concentrations, at ordinary temperature. One of these compounds is probably $CaO \cdot SiO_2 \cdot aq$. Whether there are others, and what their compositions may be, are matters that seem to be less well established.

IX. INVARIANT CONCENTRATIONS

Owing to the difficulties inherent in the investigation of a colloidal system, it is not surprising that a number of investigators have been content to treat as essentially invariant, and indicative of a phase transition, solutions that show changes of a few hundredths of a gram of CaO per liter during the rapid rise in the CaO/SiO_2 value for the solids. Changes in the SiO_2 concentration have also been disregarded. Sometimes the SiO_2 concentrations were not known, but investigators who determined them found marked variations at the assumed invariant point. Generally the SiO_2 concentration reached a maximum at about the start of the abrupt increase in CaO/SiO_2 in solids and then declined markedly during the completion of the rise. Disregard of such differences is traceable in some cases (7,88) to a belief that a part of the silica may have been in sol form. Not all authors disregarded the variations in either CaO or SiO_2; Shaw and MacIntire (158) and Flint and Wells (47) attached considerable importance to them.

In table 4 are assembled data obtained by the various investigators for the regions of abrupt change in the CaO/SiO_2 ratio in solids. The concentration ranges are given as maximum and minimum rather than initial and final values, since the maximum, especially for CaO, often occurs between the latter values. In such cases the final value is the one taken as minimum, since the initial value for the rise is commonly less definite. Maximum values are given first if they exceed the final values but not if they *are* the final values. The SiO_2 values corresponding to given CaO values are not in general brought out by this arrangement, but the minimum values are mates in all cases where reported, except for

TABLE 4

Concentrations of CaO *and* SiO_2 *in solution in regions where the* CaO/SiO_2 *value for the solid changes abruptly*

Some of the values reported were taken from curves; included are some concentration data for apparent invariant points but for which the corresponding CaO/SiO_2 values for the solids are not available

CURVE NO.	INVESTIGATORS	SOLID	SOLUTION		REMARKS
		CaO/SiO_2	CaO	SiO_2	
		molar	*grams per liter*	*grams per liter*	
1.....	Baylis (7)	0.2–1.08	0.12–0.08	0.22–0.05	
2.....	van der Burgh (28)	0.1–0.85	0.118–0.084		
3.....	Shaw and MacIntire (158)	0.46–0.80	0.12–0.07	0.23–0.03	
4.....	Flint and Wells (47)	0.1–1.07	0.102–0.049	0.33–0.05	
5.....	Jolibois and Chassevent (74)	0.83–1.53 1.17–1.53	0.033–0.053 0.055–0.053	0.22–0.06	
6.....	Beitlich (8)....	0.28–0.77 0.28–0.53	0.041–0.073 0.041–0.042		
7.....	Krasil'nikov and Kiselev (88)	0.1–0.9	0.090 0.086 0.087	0.20–0.03	The CaO values of 0.086 and 0.087 were obtained in experiments on tripoli and diatomaceous earth
8.....	Bessey (11)	–0.97	0.050–0.055	0.035–0.050	Concentration ranges reported by author without comment
9.....	Roller and Ervin (152)	0.82–0.87 0.97–1.08	0.050 0.173	0.041+ 0.006	The SiO_2 concentration of 0.041 g./l. was obtained at 0.056 g. CaO per liter, the lowest CaO concentration tested
12.....	Kühl and Mann (97)	0.1–1.0	0.11–0.13		
13.....	Thorvaldson and Vigfusson (178)	0.16–1.07	0.06–0.15		This line is for $2CaO \cdot SiO_2$
		0.37–1.10	0.19–0.24		This line is for $3CaO \cdot SiO_2$ (see Section XVI)
14.....	Cirilli (34, 35)	0.53–1.00	0.065–0.075		
15.....	Tavasci (167)	0.2–1.44	0.03–0.077	0.02–0.038	
	Le Chatelier (117)		0.052		
	Michaelis (134)	(0.5)	0.051	0.042	Results of a single test when CaO/SiO_2 in solids was about 0.5

TABLE 4—*Concluded*

CURVE NO.	INVESTIGATORS	SOLID	SOLUTION		REMARKS
		CaO/SiO_2	CaO	SiO_2	
		molar	*grams per liter*	*grams per liter*	
	Lafuma (104)		0.052		This invariant concentration was reported in work on portland cement
	Kolthoff and Stenger (86)		Less than 0.028		See discussion at end of reference 86
	Maffei (126)		0.09		The data suggest that the value is between 0.09 and 0.13; K. and K. (88) give it as 0.09
	Cummins and Miller (36)		0.11	0.18	pH given as 9.6; solid phases indicated were silica and $5CaO \cdot 6SiO_2 \cdot aq$
	Weisberg		0.046		Reported by Bessey (11); no reference given to previous publication
	Kind and Aleksandrovskiĭ (83)		0.118		Cited by Krasil'nikov and Kiselev (88)

Jolibois and Chassevent's tests and perhaps in Bessey's tests. Bessey reported his ranges without comment, and the statements made here do not necessarily apply in his case.

Where a concentration range is not reported this does not necessarily mean that the absence of one was demonstrated. In some of these cases only single determinations were available. However, the separate CaO concentrations reported for Krasil'nikov and Kiselev (88) and Roller and Ervin (152) can be considered to apply over the range of solids indicated. The range in CaO shown for van der Burgh (28) is perhaps best regarded as representing the precision rather than as denoting a definite trend.

When both the CaO and SiO_2 concentrations are considered, it does not appear that any of the investigators demonstrated strict invariance.

For comparison with the values given in table 4 it may be remarked that various authors (7, 47, 152, 158) have found that the SiO_2 concentration is already reduced to less than 0.01 g. per liter when the CaO concentration has reached 0.15 g. per liter. The SiO_2 concentration continues to diminish, of course, as the CaO concentration is further increased.

As has been mentioned, Flint and Wells (47) treated their variations in CaO and SiO_2 concentrations as significant throughout. As shown in the previous section, they regarded each apparent cusp in the solubility curve as an invariant point representing a solid-phase transition. The evidence for this interpretation would have been much stronger had solids of varied CaO/SiO_2 ratios been found in apparent equilibrium with the solution represented at such a cusp. However, the preparations that were tried did not provide evidence of this nature. Krasil'nikov and Kiselev (88) pointed out that, as plotted in figure 1, Flint and Wells' data appear much like their own. The Russian authors obtained a CaO concentration so invariant that they attached no importance to the variation in SiO_2, which could be attributed to SiO_2 in sol form. Kryagova (90) has obtained results which he considers to be in essential agreement with those of Flint and Wells. However, Kryagova's comparison was made entirely on the basis of CaO/SiO_2 values, for both solutions and solids. When the individual concentrations are considered, and also when the comparison is carried to higher CaO/SiO_2 values for the solids, marked differences appear.

Relative to the retrogression in CaO concentrations obtained in several investigations as the CaO/SiO_2 ratio in solids approached 1, it is of interest that in two of these cases (Baylis (7), and Shaw and MacIntire (158)) the authors determined the pH as well as CaO and found that the former increased without significant retrogression as CaO/SiO_2 for the solids was increased. Baylis found pH values of 10.1 to 10.8 over the range in CaO/SiO_2 values of 0.1 to 1.1 for the solid. Shaw and MacIntire reported pH values a few tenths higher.

X. IONS IN SOLUTION: WORK OF FLINT AND WELLS AND OF ROLLER AND ERVIN

Owing to the difficulties presented by the $CaO-SiO_2-H_2O$ system, a knowledge of the silicate ions in solution has been sought by some of the investigators. Flint and Wells (47) sought this information for whatever clue it might give relative to compounds present in the solids. Roller and Ervin (152) estimated ion activities in order to apply the activity solubility product principle in their tests for particular solid compounds. In both cases the investigators worked directly with calcium silicate solutions. Previous studies of the silicate ions in solution have generally been made with alkali silicates. These other studies still leave much uncertainty but are commonly considered to indicate at least two stages of ionization, of either meta- or ortho-silicic acid (152) though a disilicate ion has also been assumed (72, 182). In recent articles (30, 147) the orthosilicate formulation is preferred over the metasilicate, because silicon is invariably found at the center of an oxygen tetrahedron in crystalline silicates (43) as well as in vitreous silica and soda-silica glasses (58). At the higher alkalinities (pH 10.7 and greater) apparently little or none of the silica in solution is in a colloidal state (153).

Flint and Wells concluded that the silicate ions were those of orthosilicic acid, H_4SiO_4, and that each of the four steps of acid ionization was dominant over a part of the range in $Ca(OH)_2$ concentration. They calculated each hydrolysis constant independently, treating as negligible the concentrations of silicate ions not directly involved. The final choice of concentration range for the calculation of each constant rested in part on the ratios of silicate CaO[8] to total SiO_2, and in part on the degree of invariance shown by the resultant "constants" within the trial range. Ions of orthosilicic acid, rather than metasilicic acid, were assumed by necessity, since the experimental ratio of silicate CaO to SiO_2 was greater than 1 at the higher CaO concentrations.

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[8] The CaO in solution in excess of that equivalent to OH^- was called combined CaO by the authors and is here called silicate CaO. Concentrations of CaO, SiO_2, and OH^- were obtained experimentally.

The second hydrolysis constant came out a little larger than the first, but this anomaly was attributed simply to experimental inaccuracy. Acid ionization constants calculated from the third and fourth hydrolysis constants agreed satisfactorily with ones reported by other investigators, on the basis of electrometric titrations of sodium silicate solutions. The other ionization constants could not be similarly compared, since only two steps of acid ionization were assumed in these other investigations. On the whole it appeared that the calculations gave reasonable results, and, since the various constants had been found not to change much when calculated for different solutions, the authors apparently felt justified in their assumption that the silicate ions were those of H_4SiO_4.

The results of the study can be used, however, to show that in some cases neglected ions may have been present in considerable concentrations. For example, the third hydrolysis constant, K_{h_3}, can be used to check the appropriateness of the neglect of $H_3SiO_4^-$ in the determination of the second constant, K_{h_2}. When this check is applied to the data of test 1 of Flint and Wells' table 5, $[H_3SiO_4^-] \div [H_2SiO_4^{--}]$ is indicated to be 0.57. The neglected ion is thus indicated to have a concentration of more than half that of one of the ions directly involved. Owing to indications of this kind, the present writer has calculated the four hydrolysis constants simultaneously without neglecting any of the silicate-ion concentrations. Data from the four pH ranges utilized by Flint and Wells were employed, and the constants were calculated both without and with the use of activity coefficients. In the first case the practice corresponds to that of Flint and Wells. For the second calculation the activity coefficients were estimated by treating all anions as singly charged when computing ionic strength. Comparison with more precise methods indicated that this approximation, which was also used by Roller and Ervin, is reasonably satisfactory. Irrespective of the use of activity coefficients the third and fourth hydrolysis constants came out much the same as those derived by Flint and Wells, but the first and second hydrolysis constants were negative, indicating an unreal situation.[9] To eliminate the influence of this

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[9] The four equations for the four steps of hydrolysis of Ca_2SiO_4, the equation expressing the distribution of SiO_2, and the equation for balance of electric charges were solved simultaneously to express the concentrations of silicate ions and undissociated H_4SiO_4 in terms of the hydrolysis constants. The equation obtained was

$$\frac{A}{f_0} + \frac{(OH^-)}{f_1}(A - \underline{SiO_2})I + \frac{(OH^-)^2}{f_2}(A - 2\underline{SiO_2})II + \frac{(OH^-)^3}{f_3}(A - 3\underline{SiO_2})III + \frac{(OH^-)^4}{f_4}(A - 4\underline{SiO_2})IIII = 0 \quad (1)$$

where $\underline{SiO_2}$ is the total concentration of SiO_2 in moles per liter, (OH^-) is the activity of OH^- on the molar basis, f_0 is the activity coefficient (taken as unity) for the undissociated H_4SiO_4, and the other f's are the activity coefficients for silicate ions having the number anomalous result, similar simultaneous solutions were made using only the two lowest pH ranges and assuming only two steps of hydrolysis. This also gave values comparable with those found by Flint and Wells for the third and fourth hydrolysis steps.

Thus, there may well be two stages of hydrolysis corresponding to two stages of acid ionization, but the only support that remains for the assumption that other stages of ionization are involved is the experimental indication that in some solutions the ratio of silicate CaO to SiO_2 exceeds 1. This experimental evidence must be considered in the light of the fact that the solutions concerned were very highly supersaturated with calcium silicate, and actually precipitated before the potential readings that gave the $Ca(OH)_2$ concentration were obtained. Though the authors were able to show that readings obtained at lower lime concentrations were unaffected by such precipitation, the evidence is obviously not as strong as it would have been had the solutions been stable. It should be remarked, however, that supersaturated solutions were also obtained by the action of water on anhydrous β-$2CaO \cdot SiO_2$ and $3CaO \cdot SiO_2$. It was possible to test these solu-

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of charges given by the subscripts. The A is

$$2\underline{\mathrm{CaO}} + \frac{K_w}{f_{1+}\,(\mathrm{OH}^-)} - \frac{(\mathrm{OH}^-)}{f_{1-}}$$

where $\underline{\mathrm{CaO}}$ is the total concentration of CaO in moles per liter, K_w is the ionization constan for water, and the f_{1-} and f_{1+} are the activity coefficients for OH^- and H^+, respectively The $\frac{K_w}{f_{1+}\,(\mathrm{OH}^-)}$ is the concentration of H^+ and was shown to be negligible in all instances The I, II, III, and IIII are reciprocals of K_{h_4}, $K_{h_3}K_{h_4}$, $K_{h_2}K_{h_3}K_{h_4}$, and $K_{h_1}K_{h_2}K_{h_3}K_{h_4}$ respectively, where the K's with their subscripts denote the four successive hydrolysis constants, with K_{h_1} representing the first step of hydrolysis, namely,

$$K_{h_1} = \frac{(\mathrm{HSiO_4}^{---})\,(\mathrm{OH}^-)}{(\mathrm{SiO_4}^{----})}$$

The hydrolysis constants themselves are evaluated as follows:

$$K_{h_4} = \frac{1}{\mathrm{I}},\ K_{h_3} = \frac{\mathrm{I}}{\mathrm{II}},\ K_{h_2} = \frac{\mathrm{II}}{\mathrm{III}},\ K_{h_1} = \frac{\mathrm{III}}{\mathrm{IIII}}$$

Equation 1 was set up for four different solutions, one for each of Flint and Wells' pH ranges within which $\underline{\mathrm{SiO_2}}$, $\underline{\mathrm{CaO}}$, and $\frac{(\mathrm{OH}^-)}{f_{1-}}$ had been determined experimentally. From the four resultant equations the K_h values were determined by simultaneous solution. Using data for which the Flint and Wells table and test numbers are 3:8, 4:8, 5:7, and 6:1, the writer obtained the following results:

	ALL ACTIVITY COEFFICIENTS TAKEN AS UNITY	ACTIVITY COEFFICIENTS ESTIMATED FROM $-\log_{10}f = \frac{0.505z^2\sqrt{\mu}}{1 + z\sqrt{\mu}}$ (82)
K_{h_4}	7.0×10^{-5}	7.2×10^{-5}
K_{h_3}	4.7×10^{-3}	8.6×10^{-3}
K_{h_2}	-1.5×10^{-2}	-6.2×10^{-2}
K_{h_1}	-1.2×10^{-2}	-1.0×10^{-2}

tions before precipitation (190), and they also showed ratios of silicate CaO to SiO_2 that were greater than 1. Such evidence is obviously better, though the instability of the solutions is still a disturbing factor. One can imagine that the very strongly supersaturated solutions may have contained molecular aggregations too small for detection by the available means. It appears that precise determinations made on solutions that are not supersaturated are needed to settle these questions. Should a CaO/SiO_2 value greater than 1 be established, it might mean either that more than two stages of ionization of orthosilicic acid are involved or that a complex ion is formed.

Roller and Ervin's (152) study, based on the solubility product principle, proceeds from the following reasoning: If adsorption alone is responsible for the continuous increase in CaO/SiO_2 in the solids with increase in the CaO in solution, then the adsorbing compounds, being of constant composition, should maintain constant activity solubility products over their respective regions of stability.

In their first calculations Roller and Ervin simply assumed two stages of hydrolysis corresponding to two stages of acid ionization and used ionization constants that are approximate averages of published values. For various reasons they believed that the silicate ions were derived from orthosilicic acid. Hence they assumed the ions $H_2SiO_4^{--}$ and $H_3SiO_4^-$ although, as they point out, the analysis would still be the same if SiO_3^{--} and $HSiO_3^-$ were assumed instead. They did not have direct evidence of their own for the conclusion that the $H_2SiO_4^{--}$ does not dissociate significantly, for they did not determine experimentally the ratios of silicate CaO to SiO_2. The total CaO and SiO_2 concentrations were determined directly, but OH^- was estimated. This was feasible because the SiO_2 concentration was generally very small in relation to the total CaO concentration. Estimates were compared with acid titrations carried to the phenolphthalein end point.

Proceeding on the stated basis, the authors were able to show that the compound $CaH_2SiO_4 \cdot aq$ (or $CaO \cdot SiO_2 \cdot aq$) appears to maintain a constant activity solubility product over the range in CaO concentrations that extends to the right of the final, short vertical rise in their curve (No. 9) of figure 1. However, over the range in CaO to the left of this point, constancy could only be obtained by assuming an addition compound containing about eight of the doubly charged silicate ions to each of the singly charged ones. Believing such a compound to be unreasonable, the authors preferred to assume association of the silicate ions in solution, particularly in the region of the lower CaO concentrations, with production of dimers corresponding to the acid $H_4Si_2O_6$. Proceeding on this basis, they completely redetermined all anion activities, though again they simply assumed values for certain of the ionization constants on the basis of other published data (electrometric titrations). On this new basis, the compound $CaH_2SiO_4 \cdot aq$ still proved satisfactory in the higher CaO range where the dimer ions represented only a minor part of the SiO_2. Over the lower CaO range the compound $Ca_3(HSi_2O_6)_2 \cdot aq$ (or $3CaO \cdot 4SiO_2 \cdot aq$) maintained a constant ion product, and hence appeared to be the stable compound in that region.

By these studies Roller and Ervin appear to have added materially to the grounds for the adsorption hypothesis. They make the existence of certain compounds and ions appear plausible, but the ionization constants of silicic acid and the various activity coefficients can scarcely be said to be known well enough for any more definite conclusions. The present writer has found, by trial, that a somewhat different but apparently equally appropriate way of estimating the activity coefficients $\left(\text{namely, by use of the equation} -\log_{10} f = \frac{0.505z^2\sqrt{\mu}}{1+z\sqrt{\mu}} \text{ (82)}\right)$ is sufficient to introduce a considerable progressive change in the ion product for $CaH_2SiO_4 \cdot aq$. Yet, complete constancy can be restored by making only such changes in the assumed acid ionization constants as are permitted by the range in the published values for the constants.

It does not appear that the silicate ions in solution have as yet been well established.

XI. ADSORPTION *versus* SOLID SOLUTION

The continuous increase in CaO/SiO_2 for the solids with increase in CaO in solution that occurs after the apparent formation of compounds at low CaO concentrations has been explained generally as denoting either adsorption or solid solution. If colloidal equilibrium is assumed to have been established, these are the only very obvious possibilities. According to the phase rule, the solution phase in a three-component system at constant temperature can change only when just one solid phase is present.[10] As brought out in Section III, some change is possible when two solid phases are present if the solid particles are submicroscopic in size, owing to the fact that the solubility of small particles varies with their sizes. However, it seems unlikely that such wide ranges in CaO concentration as are here concerned could be explained on this basis, or that, if the solid phases *are* of fixed composition and non-adsorptive, horizontal curve sections representing single solid phases should not appear in addition to the sloping sections (figure 1).

Several investigators followed Le Chatelier in assuming adsorption. Baylis (7), Shaw and MacIntire (158), Beitlich (8), and Roller and Ervin (152) reported that the data for their sloping curve sections conformed to Freundlich adsorption isotherms. However, only Roller and Ervin appear to have taken due account of the chemically bound CaO, by deducting the CaO in the assumed compound. In any case, agreement with a Freundlich isotherm is simply an indication rather than a proof of adsorption; also, disagreement does not disprove adsorption (59, 187). The use made of the solubility product principle by Roller and Ervin provides further evidence in support of adsorption but, as previously discussed, is also not decisive. The fact that Chassevent (31) obtained the same x-ray spectrum for gels of varied lime content is interesting, but the author himself remarked that the bands were broad and not sharply enough defined to eliminate the possibility of some solid solution. Baylis (7) made an observation that seems to warrant further investigation; he made silica gels with varied amounts of water and found an apparent influence on the amount of CaO taken up, the denser gels taking up the least CaO.

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[10] Adsorbed material is not to be treated as an additional phase in applications of the phase rule (4).

Little relative to ion-exchange properties has been reported. However, in the presence of sodium hydroxide in the lime solution, the calcium silicate gel loses some of its CaO, which is replaced by a lesser amount of Na_2O (see Section XVII). Sloane, McCaughey, Foster, and Shreve (160) found that when β-$2CaO \cdot SiO_2$ that had been placed in a solution of calcium chloride was washed first with water and then with magnesium nitrate solution the first portion of magnesium nitrate solution increased markedly the rate at which lime was brought into solution. A similar test was later made by Steinour and Ledyard (163) on β-$2CaO \cdot SiO_2$ hydrating in plain water. Use of magnesium chloride solution increased the rate at which lime was brought into solution, though in this case no calcium chloride had been employed. To be considered, of course, in both cases, is the extent to which the results may have been influenced by double decomposition. However, Steinour (162) found, by electroösmosis, that β-$2CaO \cdot SiO_2$ particles in plain water, and in $Ca(OH)_2$ solution, carry a positive charge which can be varied by solutions of salts in accordance with the Schulze–Hardy rule, and can be reversed by sufficiently high concentrations of salts having negative ions of higher valence than 1. Calcium and barium salts having monovalent anions were found to increase the positive charge. In this connection it is of interest that Shaw and MacIntire (158) found that calcium chloride increased the CaO/SiO_2 value of the lime–silica gel.

Of the various investigators who assumed solid solution, Cirilli (32, 34) appears to have been the only one to advance experimental evidence for this preference. He found, as has already been reported, that the heat of solution of the gel, per gram of CaO, remained constant for gels for which the CaO/SiO_2 value was less than 1; but when CaO/SiO_2 was greater than 1, the increase in heat of solution was more than it should have been for $CaO \cdot SiO_2 \cdot aq$ plus $Ca(OH)_2$. However, though the difference reached a magnitude of about 20 cal. per gram of CaO at $CaO/SiO_2 = 1.5$, the present writer is more impressed by the degree of correspondence than by the differences, especially since there is no definite indication of a break in the curve as it is continued into the region beyond $CaO/SiO_2 = 1.5$, where Cirilli (35) found free $Ca(OH)_2$. Possibly the thermal discrepancy found by Cirilli can be accounted for on the basis of the fine subdivision of the adsorbed lime.

Bessey (11) and Cirilli (34) thought it probable that solid solutions are formed between $CaO \cdot SiO_2 \cdot aq$ and $3CaO \cdot 2SiO_2 \cdot aq$, because at the end of the gradual rise in CaO/SiO_2 this ratio was about 1.5. However, as Keevil and Thorvaldson (80) remarked, there is no very substantial evidence for assuming that the gel for which CaO/SiO_2 is 1.5 is the definite compound $3CaO \cdot 2SiO_2 \cdot aq$. The solid solution might be between $CaO \cdot SiO_2 \cdot aq$ and $2CaO \cdot SiO_2 \cdot aq$; the value of CaO/SiO_2 that happens to be reached when the solution becomes saturated with $Ca(OH)_2$ is not very informative (101). Bessey, of course, was influenced by the belief that not $Ca(OH)_2$, but $2CaO \cdot SiO_2 \cdot aq$, was formed next after the gradual rise in the curve.

Experimental evidence bearing upon the kind of solid solution that might be formed is meager. Kühl and Mann (102) regarded the gel as initially amorphous, evidently on the basis of negative results of early x-ray diffraction tests (92). Bessey (11) also wrote of the gel as amorphous. However, the latest x-ray evidence indicates a finely crystalline structure (Section VII). Büssem (25) presents a brief discussion in which he seems to imply mixed crystals, or solid solution based on supposed isomorphism of $CaO \cdot SiO_2 \cdot aq$ and $3CaO \cdot 2SiO_2 \cdot aq$. He mentions in this connection the fact that Chassevent, and Strätling, obtained practically the same x-ray spectrum, though they believed their gels to be $CaO \cdot SiO_2 \cdot aq$ and $3CaO \cdot 2SiO_2 \cdot aq$, respectively.

Lafuma (105), Tavasci (167), and Keevil and Thorvaldson (80) have suggested that the gel is zeolitic with respect to its variability in lime content. Though the term zeolitic is sometimes used very broadly (42), Lafuma makes it clear that he has in mind an open crystalline structure like that of the true zeolites. The expansible layer structure of colloidal clays, called zeolitic by some authors (42), can also be regarded as a possibility. In either case, the structures can evidently be regarded as solid solutions, since they constitute single phases.

A fact that needs to be borne in mind when considering these possibilities is that the solid can apparently vary continuously in equilibrium water content much as it does in CaO content but independently of the latter variation (see Section XV). It appears, also, that for some preparations the time required for complete readjustment in CaO or H_2O is no more than a few hours or a few days. Either an open or expansible structure or an exceedingly fine state of subdivision seems to be indicated. Indeed, if the gel is $CaO \cdot SiO_2 \cdot aq$ plus adsorbed lime, it would seem that the degree of subdivision of the solid must approach the molecular in one or more dimensions, because 0.5 to 1 mole of $Ca(OH)_2$ must be adsorbed per mole of $CaO \cdot SiO_2 \cdot aq$. In this connection, Roller and Ervin (152) have pointed out that Weiser, Milligan, and Bates (188) have apparently demonstrated similar degrees of adsorption by other compounds. An adsorptive or a zeolitic type of structure is suggested by x-ray studies made by Strätling (165). This investigator found that heating the colloidal hydrous calcium silicate to 250°C. did not change the x-ray spectrum. Also, Mlle. Foret (49) reported that variation in the water contents of her hydrothermally prepared crystalline monocalcium silicate hydrates did not change their x-ray spectra.

It seems worth mentioning that zeolitic or any other type of solid solution does not explain, as readily as does adsorption, the different CaO/SiO_2 values obtained in various supposed equilibrium studies at the same CaO concentrations. If the internal structures are the same, the CaO/SiO_2 values at equilibrium should apparently be the same also, except as they are influenced by different amounts of external adsorption. It seems possible, however, that an expansible layer structure might expand irregularly (68), with dependence upon the manner in which it had been prepared or upon traces of impurities. In any case, the fact should not be overlooked that a zeolitic solid solution can evidently be considered a special kind of internal adsorptive system. Indeed, synthetic ion-exchange resins, of zeolitic nature, are commonly called exchange-adsorbents, and provide data that conform to the Freundlich adsorption isotherm (136). When a zeolitic structure adsorbs physically simply by reason of its fine porosity, the process resembles physical adsorption (5); on the other hand, base exchange in a zeolite is evidently closely related to chemical adsorption.

XII. PRODUCTS FORMED IN SATURATED AND SUPERSATURATED LIMEWATER

Table 5 shows the CaO/SiO_2 values for the solids that were found by various authors to be in apparent equilibrium with solutions saturated or nearly saturated with $Ca(OH)_2$. Some values have also been added, with a plus sign, which were obtained in investigations in which the highest concentration of $Ca(OH)_2$ in solution was still considerably below the saturation value. Only studies in which the volume of solution was very large in proportion to the weight of solid are reported; studies made on pastes are reported in the next section. The states of chemical combination that were assumed by the various investigators are given in the final column of the table. Where the theoretical ratio is less than the experimental one, adsorption was ordinarily assumed. Klasse and Kühl (84) and Flint and Wells (47) obtained values distinctly below 2, but they were inclined to assume that, had equilibrium been fully attained, $2CaO \cdot SiO_2 \cdot aq$ might have resulted.

There is not, however, the same reason to assume some definite silicate compound at the lime saturation point as there is at the top of the nearly vertical curve section (figure 1) that occurs at low CaO concentration. The change from a gradual slope to a nearly vertical one that occurs at lime saturation means simply that $Ca(OH)_2$ begins to precipitate. Kühl and Mann (97) found that when the solution supersaturates with respect to $Ca(OH)_2$, the curve (No. 12, figure 1) continues smoothly beyond the saturation point. The rise becomes more rapid as the saturation point is passed, but the authors were inclined to attribute this to deposition of some $Ca(OH)_2$. Hedin (67), however, obtained a similar shape of supersaturation curve (No. 10, figure 1) and was unable to detect free $Ca(OH)_2$ microscopically. As $Ca(OH)_2$ might be deposited in a very fine state, the inability to observe it probably does not demonstrate its absence. Of some interest in this connection is the fact that Thorvaldson and Vigfusson (178) found that their curves (No. 13, figure 1) passed through the lime saturation point without a significant change in slope. In the investigations of Kühl and Mann, and of Hedin, CaO/SiO_2 values of 2 were reached and the authors were of the opinion that much of the controversy in the literature as to the CaO/SiO_2 value obtainable in the gel is the result of varied degrees of supersaturation. The ability of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ to form supersaturated $Ca(OH)_2$ solutions which can remain supersaturated for long periods is well known (87, 103, 122, 137, 167). Hedin advanced the idea that if $2CaO \cdot SiO_2 \cdot aq$ has begun to form as a result of supersaturation of the solution with $Ca(OH)_2$, it may perhaps still continue to form after the concentration of the solution has been lowered.

Forsén (52) reported experiments which led him to conclude that the hydrated calcium silicate in equilibrium with saturated limewater is $2CaO \cdot SiO_2 \cdot 4H_2O$.

TABLE 5

*Composition of lime-silica gel found in apparent equilibrium with saturated limewater**

CURVE NO.	INVESTIGATORS	MOLAR CaO/SiO_2 IN SOLIDS	INTERPRETATION GIVEN
1	Baylis (7)	1.2–1.5	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime
2	van der Burgh (28)	1.4	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime
3	Shaw and MacIntire (158)	1.3	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime
4	Flint and Wells (47)	1.85	No definite conclusion, but $2CaO\cdot SiO_2\cdot aq$ suggested
5	Jolibois and Chassevent (74)	1.6+	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime (see also reference 31)
6	Beitlich (8)	2.0	$2CaO\cdot SiO_2\cdot aq$
7	Krasil'nikov and Kiselev (88)	1.9+	No definite conclusion
8	Bessey (11)	2.1	$2CaO\cdot SiO_2\cdot aq$
9	Roller and Ervin (152)	1.45	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime
10	Hedin (67)	1.5	$3CaO\cdot 2SiO_2\cdot aq$
11 12	Kühl and Mann (97)	1.5	No definite conclusion
13	Thorvaldson and Vigfusson (178)	1.45	No definite conclusion
14	Cirilli (34)	1.4	$3CaO\cdot 2SiO_2\cdot aq$ (Cirilli called the saturation ratio 1.5)
15	Tavasci (167)	1.7	$5CaO\cdot 3SiO_2\cdot aq$
16	Nacken (137)	2.15+	$2CaO\cdot SiO_2\cdot aq$ (in reference 137; in reference 138 a marked decrease in ratio was reported)
	Keevil and Thorvaldson (80)	1.5	No definite conclusion
	Le Chatelier (117)	1.7	$CaO\cdot SiO_2\cdot aq$ plus adsorbed lime
	Klasse and Kühl (84)	1.8	$2CaO\cdot SiO_2\cdot aq$
	Kühl and Wang Tao (103)	1.5†	$3CaO\cdot 2SiO_2\cdot aq$
	Koyanagi (87)	1.0†	$CaO\cdot SiO_2\cdot aq$
	Michaelis (132, 133)	1.5	At first M. assumed $3CaO\cdot 2SiO_2\cdot aq$; later he drew no definite conclusions
	Maffei and Battaglia (127)	1.25	No definite conclusion (value given is taken from M. and B.'s table 1, data for silica sol)
	Landrin (107)	1.33‡	L. appears to assume $4CaO\cdot 3SiO_2\cdot aq$

* A plus sign indicates that the CaO/SiO_2 value shown was the highest obtained in an investigation in which the highest concentration of $Ca(OH)_2$ was still far short of saturation. Shaw and MacIntire's value was obtained when the solution was nearly saturated with $Ca(OH)_2$.

† Experiments on portland cement clinker (corrected).

‡ From reference 107 it is not wholly clear whether the final solution was completely saturated with $Ca(OH)_2$. However, the ratio 1.33 was considered to represent the maximum possible absorption of lime.

Solutions obtained by treating $3CaO\cdot SiO_2$ with water were poured into saturated $Ca(OH)_2$ solutions (proportions of 1:4 by volume). Precipitates were formed which after drying at 105°C. had in each case approximately the composition

$2CaO \cdot SiO_2 \cdot 4H_2O$. On the other hand, Hedin (67), whose initial studies were made in collaboration with Forsén, reports precipitations of similar character in which the values of CaO/SiO_2 attained in the products were not as high. The $3CaO \cdot SiO_2$ extracts and the saturated $Ca(OH)_2$ solutions were mixed in the proportions of 1:1 and 1:5. The precipitates, removed after 24 hr., showed on analysis the compositions $1.6CaO:1SiO_2:aq$ and $1.8CaO:1SiO_2:aq$, respectively.

The investigations of Kühl and Wang Tao (103), and of Koyanagi (87), cited in table 5, fall in a special category since they are not strictly $CaO–SiO_2–H_2O$ studies but were made by use of portland cement clinker, with correction for hydration products containing Al_2O_3 and Fe_2O_3. The results are of special interest, however, because they are distinctly different for the two investigations, and this appears to be attributable largely, though not entirely, to one circumstance. That is, Kühl and Wang relied on physical separation of $Ca(OH)_2$ crystals, whereas Koyanagi supplemented this procedure by testing and correcting for free $Ca(OH)_2$ retained by the gelatinous precipitate. The corrections for Fe_2O_3 were not the same in the two investigations, but the effect of the difference was small and unimportant. The experimental procedure consisted in shaking grains of the clinker with saturated limewater. Voluminous hydration products separated from the grains and were separately analyzed. When the experiments were continued for a long time, coarse $Ca(OH)_2$ crystals formed and settled out.

As shown by table 5, Koyanagi's corrected CaO/SiO_2 value was 1, in agreement with adsorption theories, whereas Kühl and Wang obtained 1.5, in agreement with the gross compositions reported by several other investigators. The test for free $Ca(OH)_2$ used by Koyanagi was the glycerol–alcohol solution method, with titration with ammonium acetate (123). As used for *hydrated* lime the test has given difficulty, apparently because of inadequate solution of the crystals of $Ca(OH)_2$, with resultant low values for $Ca(OH)_2$ (15, 18). However, such an effect is in the wrong direction to account for the difference in the results obtained by Kühl and Wang and by Koyanagi. As a matter of fact, Koyanagi's final values of molar CaO/SiO_2, even before correction for the free $Ca(OH)_2$, were distinctly less than those of Kühl and Wang, being only about 1.3. Probably a little alkali hydroxide was present in solution in both investigations, though the concentrations must have been small since Kühl and Wang, and Koyanagi, used 10 and 20 parts of limewater to 1 part of clinker, respectively. (For effects of alkalies see Section XVII.)

Bessey (11) has criticized Koyanagi's conclusions on the ground that "lime which may be present in excess of a 1:1 lime:silica ratio is apparently in a weak state of combination and is likely to react with glycerol or other reagents in a manner similar to free lime." In this connection it is of interest that Hedin (67) used ethylene glycol to extract free $Ca(OH)_2$ from a product obtained by reacting sodium silicate with a suspension of $Ca(OH)_2$. The extracted product had a CaO/SiO_2 value of 1.5. Like Bessey, Hedin was fearful that the reagent could extract chemically combined lime, so he abandoned the procedure. However, when he used tetraethyl silicate with saturated limewater in the absence of solid $Ca(OH)_2$ the product, though not extracted, also had a CaO/SiO_2 value of 1.5. This seems to indicate that the ethylene glycol was not too aggressive a solvent, and no reason is seen why glycerol would be more active. However, the fact remains that Koyanagi found CaO/SiO_2 equal to 1.0, whereas Hedin found 1.5. It may at least be noted that Hedin dried his product at 105°C. before extraction, whereas Koyanagi apparently dried his at room temperature (*in vacuo*), though it was heated later during the free lime test.

Koyanagi found that for test periods of only 7–10 days the corrected CaO/SiO_2 value was about 1.2–1.3, but that as the test periods were extended the value dropped to 1.0 and remained there. From the higher initial ratio Koyanagi assumed that a metastable silicate higher limed than $1CaO:1SiO_2$ is first formed. This silicate could have a considerably higher CaO/SiO_2 value than 1.3, since in Koyanagi's preparations some $CaO \cdot SiO_2 \cdot aq$ would probably have been present along with it.

From this survey it is obvious that the extent of actual chemical combination attained in calcium silicate gel in equilibrium with saturated or supersaturated limewater is still obscure.

XIII. CALCIUM SILICATE PASTES AND RELATED SUBJECTS

The two calcium silicates of portland cement, $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$, can both develop saturated solutions of $Ca(OH)_2$ when placed in contact with limited amounts of water. That this is true under apparent equilibrium conditions, brought about by vigorous agitation with an attrition agent, is evident from curves 13 of figure 1. It is also true under quiescent conditions. Indeed, either compound can cause supersaturation with respect to $Ca(OH)_2$ (122), but this condition, though somewhat persistent, must nevertheless be temporary. Accordingly, when these anhydrous calcium silicates are made up with water to the consistency of portland cement pastes, the final solutions in the hardened pastes are evidently saturated solutions of $Ca(OH)_2$.

Experimental determination of the composition of the hydrous calcium silicate that is formed in pastes is beset with difficulty, because of the intimate mixture of this product with $Ca(OH)_2$ that has precipitated and with any anhydrous material that has not yet reacted. When finely ground $3CaO \cdot SiO_2$ is used, the latter difficulty is not serious because the hydration is practically complete in specimens that have been stored for several months. However, β-$2CaO \cdot SiO_2$ contains much unhydrated material even after two years (18), unless the hardened paste is repeatedly ground and remixed with water. As to the calcium hydroxide that precipitates, the attempt is generally made either to dissolve it out or to apply a quantitative analytical test for it. In early studies the methods used appear to have been generally unsatisfactory (121). Methods have been developed in recent years that are reasonably adequate for the determination of free CaO in unhydrated cement and in preparations of the clinker compounds, but neither these nor other current methods are generally accepted as being wholly satisfactory for the accurate determination of

$Ca(OH)_2$ in the hydrated products. As mentioned earlier, inaccuracy has apparently been caused in some cases by difficulty in dissolving the $Ca(OH)_2$ crystals (15, 18). Some investigators (18, 51) have heated their hydrous products to 530°C. or more in order to dehydrate the $Ca(OH)_2$ before dissolving the lime. However, in tests on six-month specimens, Work and Lasseter (191) found less free lime in preheated (850°C.) specimens than in others not heated, a result opposite to that to be expected if the test made without heating is inadequate simply because of difficultly soluble $Ca(OH)_2$. It may be that the heating sometimes causes significant recombination of lime with calcium silicate, though tests made by Work and Lasseter on specimens that had hydrated for relatively short periods did not indicate this.

Various methods of determining $Ca(OH)_2$ in hydrated cements were discussed and compared by Bessey (13, 15) at the symposium held by cement chemists and technologists at Stockholm in 1938 (141). In a calorimetric method developed by Bessey (10), the $Ca(OH)_2$ is estimated from the difference in heat of hydration of samples previously heated to 350° and 550°C., respectively, the $Ca(OH)_2$ having been found to dehydrate between these temperatures. The other methods in common use are all extraction, or lime-dissolution, methods; Bessey made comparative tests on hydrated portland cements using, besides the calorimetric method, a half-saturated lime-solution method due to Bakewell and Bessey (3), the glycerol–alcohol solution method of Emley (44) which was first used successfully with unhydrated cements by Lerch and Bogue (121, 123), and the ethylene glycol method of Schläpfer and Bukowski (26, 156). Still other methods and also modifications of some of these have been published, but it is not the purpose here to review the various procedures. It may be noted, however, that more recently a new solvent has been proposed by Franke (54), one composed of acetoacetic ester and isobutyl alcohol.

Application of such methods to the hydrated pastes of $3CaO \cdot SiO_2$ or of portland cement has given the CaO/SiO_2 values that are shown in table 6. In the case of the cements, these ratios have been corrected for CaO in aluminates and ferrites. The methods of determining the $Ca(OH)_2$ are indicated. Bessey made tests by the four different methods already mentioned but relied primarily on the results of the calorimetric method, which were indeed practically the same as those obtained with limewater. The other two methods gave rather different results, but the average CaO/SiO_2 values were still much the same.

From table 5 it might have been expected that the CaO/SiO_2 values of table 6 would have fallen principally within the range 1.4–2.0. However, if the results obtained with cements are excluded as being subject to more sources of error than the others, there remain five values between 2.0 and 2.4 and only one that is lower. Thus, it appears that tests on pastes generally give somewhat higher values than are obtained in equilibrium studies on dilute suspensions.

Several experimenters (18, 23, 173, 191) have made studies on β-$2CaO \cdot SiO_2$ pastes but evidently without complete hydration having been attained in any case. None of these investigators found much free $Ca(OH)_2$. Lerch and Bogue (18) found that after two years the β-$2CaO \cdot SiO_2$ had liberated only 0.9 per cent of its CaO, though it had "fixed" 12 per cent of water so that it was not removable by drying at 105°C. Brown and Carlson (23) found no appreciable $Ca(OH)_2$ in a β-$2CaO \cdot SiO_2$ paste that had, by repeated grinding and mixing, been hydrated to the extent of about 25 or 30 per cent of the $2CaO \cdot SiO_2$.

TABLE 6

Composition of hydrous calcium silicate as formed in pastes

INVESTIGATOR	METHOD OF DETERMINING FREE $Ca(OH)_2$	INITIAL SOLID	AMOUNT OF WATER USED	DURATION OR NATURE OF TREATMENT	CaO/SiO_2 IN HYDROUS CALCIUM SILICATE*
			per cent of initial solid		
Assarsson and Sundius (2)	Glycerol–alcohol	Portland cement			1.9
Work and Lasseter (191)	Glycerol–alcohol	$3CaO \cdot SiO_2$	45–100	3 and 6 months	1.6–1.7
Bogue and Lerch (18)	Glycerol–alcohol with preliminary dehydration of $Ca(OH)_2$	$3CaO \cdot SiO_2$	50	1 and 2 yr.	2.2
Brown and Carlson (23)	Glycerol–alcohol	$3CaO \cdot SiO_2$	75	10 days "kneaded"	2.4
Schläpfer (155) (and Berchem)	Ethylene glycol	$3CaO \cdot SiO_2$ Portland cement	45–49		2.1 1.5
Haegermann (63)	Glycerol–alcohol with preliminary dehydration of $Ca(OH)_2$	$3CaO \cdot SiO_2$		Repeated grinding and mixing with water	2.0
Bessey (13)	Calorimetric	Portland cements (six different ones)	"Thin paste"	Repeated grinding and mixing with water	1.5
Meyers (131)		$3CaO \cdot SiO_2$	"Plastic"	Several months	2.2

* After correction for free $Ca(OH)_2$.

Though Lerch and Bogue found little hydrolysis beyond $CaO/SiO_2 = 2$ in pastes, they believed from earlier study that the equilibrium composition for hydrous calcium silicate in contact with a large excess of saturated limewater corresponds to $CaO/SiO_2 = 1.5$. They therefore concluded that the pastes do not reach equilibrium. Bessey, who obtained CaO/SiO_2 values of 1.5 in his paste studies (13), believed on the basis of his other work (11) that in the system $CaO–SiO_2–H_2O$ the equilibrium CaO/SiO_2 value is 2.0 at lime saturation. He concluded, however, that 1.5 is probably an equilibrium value for cement and that it differs from 2 because of the other components in solution besides CaO and SiO_2 (14).

In pastes the gel, forming compactly and evidently to a considerable extent as coatings over the still hydrating cores of the particles, may not represent the final equilibrium product but may be conditioned by the composition of the initial solid. This may perhaps explain why the hydration products formed around the grains of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ are found, in most studies, to have relatively high values of CaO/SiO_2. It might also be supposed on similar grounds that a coating forming over hydrous silica might have a relatively low CaO/SiO_2 value. It is of interest, therefore, that van der Burgh (28), who allowed saturated limewater to act on silica gel under quiescent conditions, found that thin membranes of reaction products were formed, which in different tests had CaO/SiO_2 values of only 0.85, 0.81, 0.81, and 0.90. Recall also the test by Lerch and Bogue (122), reported in Section VII, in which a sample of hydrous silica kept in saturated limewater until 97 per cent of the silica was acid-soluble showed CaO/SiO_2 of only 1.07 in the acid-soluble part.

Since β-$2CaO \cdot SiO_2$ apparently has the orthosilicate structure with separate SiO_4^{4-} anions (20, 43), it may be that a hydration product involving these atom groupings is relatively easily formed. At the boundary between anhydrous particle-core and gel coating the scarcity of free space may make it difficult for $Ca(OH)_2$ to crystallize, and may promote local supersaturation favoring formation of a high-limed product. This might account for the meager hydrolysis of β-$2CaO \cdot SiO_2$ in pastes, as compared with the hydrolysis obtained in such studies as that of Thorvaldson and Vigfusson (178) in which the conditions for readjustment of the initial gel composition were evidently much better.

With respect to $3CaO \cdot SiO_2$, the crystal structure of this compound may also promote the formation of a hydrated dicalcium silicate, even though that product may not be the final equilibrium one. Early theorizing from the standpoint of structural chemistry (129) led various investigators to the belief that $3CaO \cdot SiO_2$ is a basic salt in which one of the oxygen atoms is not directly attached to silicon. Though the original grounds for this deduction were inadequate (62), the conclusion itself now seems almost certainly correct. The detailed crystal structure of $3CaO \cdot SiO_2$ has not been worked out, but there is every reason to believe that the silicon–oxygen arrangement is determined by the same principles as in other silicates. That is, oxygen atoms are invariably found to form tetrahedra about the silicon atoms (43). Accordingly, $3CaO \cdot SiO_2$ is evidently composed of the ions $3Ca^{++}$, SiO_4^{4-}, and O^{--} (25, 62). The formation of $Ca(OH)_2$ from one mole of CaO per mole of $3CaO \cdot SiO_2$ is therefore not hydrolysis in the usual sense but rather a direct hydration of CaO that is released upon dissolution of the solid (102). It may be presumed that the strong tendency for this one CaO to form $Ca(OH)_2$ is at least one of the reasons why

$3CaO \cdot SiO_2$ is so much more reactive toward water than is β-$2CaO \cdot SiO_2$. As the O^{--} ion unites with water and, with Ca^{++}, forms $Ca(OH)_2$, the remainder of the $3CaO \cdot SiO_2$ is evidently enabled to hydrate also, perhaps without any immediate hydrolysis (80). Some authors (191)[11] have observed relative to the initial stages of the hydration that for each mole of water that goes to form $Ca(OH)_2$, approximately 1 mole of water is firmly bound by the silicate.

Some investigators (103) have believed that the calcium silicates can hydrate without complete solution of the whole original siliceous structure. While some experimental results do suggest this (103), there seems to be good evidence that both $3CaO \cdot SiO_2$ and $2CaO \cdot SiO_2$ dissolve in a normal manner when in the presence of ample water (52, 67). Possibly, however, under the crowded conditions at the interface between anhydrous particle and hydrous gel layer, the silicate ions do not necessarily all pass through a true solution phase during the hydration (96).

Diehl, in observations reported by Nacken (137), found that $3CaO \cdot SiO_2$ particles mounted on a gelatin base hydrated as pseudomorphs. A fringe of hydrate was observed to grow inward with the passage of time leaving, after a few days, a completely hydrated product patterned after the original particle and apparently of the same size. In some cases a second fringe was observed after a few days, which grew inward through the part already hydrated, thus appearing to indicate that the product first formed is not the final equilibrium one. The products gave a slight indication of double refraction, and an x-ray picture was obtained which seemed to indicate fine crystallinity. Baylis (7) observed that particles of precipitated calcium silicate lose lime to acid solutions and leave skeletons of hydrous silica that retain the shapes of the original particles. Thorvaldson and Shelton obtained siliceous skeletons upon treating crystalline calcium silicate hydrate with acid (174, 177).

That the bonding of hydrous reaction product onto the unreacted remainder of a calcium silicate particle may be very firm and highly effective in slowing up further hydration and hydrolysis is indicated by observations of Thorvaldson and Vigfusson (178). These investigators found, as previously mentioned, that when they did not use metal shot their solutions ceased to change in concentration while the solid particles still had unreacted cores. Also, Kühl and Wang Tao (103) found that when clinker grains are shaken with water the impacts between particles, with resultant disturbance of the hydration layers, are important in promoting further reaction. They found that the rate at which the clinker grains hydrated was distinctly less when 3 g. was shaken with 300 ml. of water or limewater than when 30 g. was used. In this connection it is worth noting that Nacken (137) used only a gram or two of $3CaO \cdot SiO_2$ per liter, without an attrition agent. The fact that less hydrolysis occurred than in some of the other studies may be traceable to this. In a few tests in which he did use an attrition agent, the hydrolysis was much increased.

Rate-of-hydration curves obtained by Thorvaldson and Vigfusson (178) for 200-mesh $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ shaken in water with metal shot are shown in figure 2. The peculiar shape of the curve for β-$2CaO \cdot SiO_2$ was said to be characteristic. Though no such irregularity was found by Lerch and Bogue (122), the fact that it was consistently obtained with one preparation of β-$2CaO \cdot SiO_2$ is still of interest. To Thorvaldson and Vigfusson it suggested that the hydrolysis takes place in two steps determined by the constitution of the molecule of β-$2CaO \cdot SiO_2$. To the present writer it suggests interference by a coating of hydrate which was difficult to dislodge. The lime liberation during the initial abrupt rise in the curve would produce a concentration of about 0.14 g. CaO per liter, evidently sufficient for formation of $CaO \cdot SiO_2 \cdot aq$. It can therefore be assumed that the β-$2CaO \cdot SiO_2$ became coated with monocalcium silicate hydrate which was in equilibrium with the solution and resisted removal. After the slowly continuing hydration had materially increased the bulk of the hydrate layer, it might have torn loose more easily.

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[11] Krauss and Jörns (89) made a similar observation, but their nominal $3CaO \cdot SiO_2$ was prepared from a melt and was therefore evidently $2CaO \cdot SiO_2$ plus CaO (37).

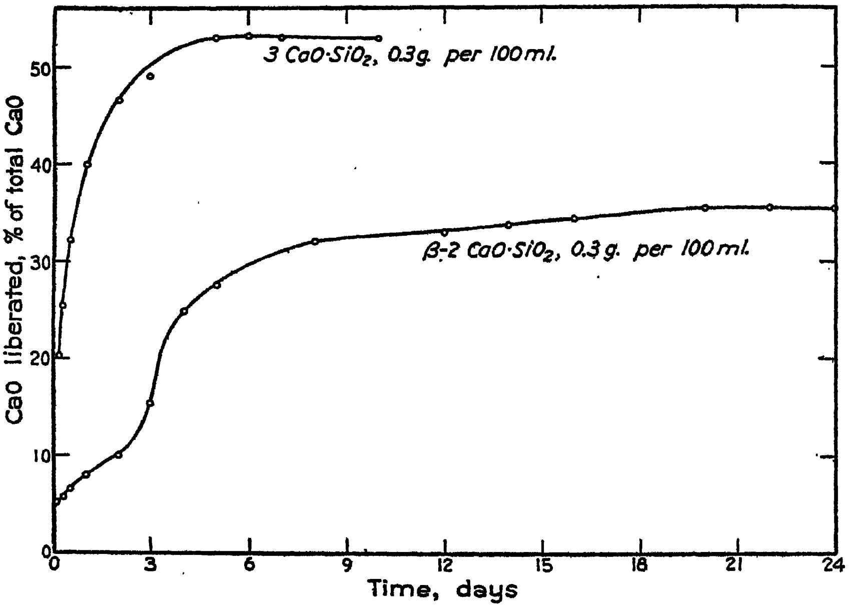

FIG. 2. Rates of hydrolysis of 200-mesh $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$; metal shot was used as attrition agent (Thorvaldson and Vigfusson (178)).

Phenomena that can be attributed to an adherent coating have also been observed in studies in which the initial solid was hydrous silica. Cummins and Miller (36), who shook diatomaceous earth with limewater and followed the reaction by conductivity measurements, observed an initial, extremely rapid reaction followed by a period of inhibited reaction which appears to have begun within an hour and continued for several hours, after which the rate again became greater. The authors attributed the inhibition to a coating of hydrated calcium silicate gel but did not explain the subsequent rise. Maffei and Battaglia (127) and Cirilli (32) reported that the rate of interaction between silica gel and limewater falls off at an early period and subsequently increases; they attributed the increase to a dispersing, or peptizing, action of the lime, as noted earlier.

The various opinions on the diminutions and subsequent increases in rate of reaction are of course largely speculative. Still another idea, advanced by Lorenz and Haegermann (124) to explain an observed slowing up and subsequent increase in rate of hydrolysis of cement grains, is that during the period of apparent marked reduction in hydrolysis lime is being absorbed by the cement particles.

XIV. MICROSCOPICAL STUDIES

Between 1908 and 1915 numerous investigators (1, 16, 55, 56, 81, 85, 108, 135, 148, 154, 164) examined microscopically the hydration of samples of portland cement clinkers or cements and in some cases direct calcium silicate preparations ("melts"). Generally, a very little of the powdered sample was placed in a drop or two of water over which a cover glass was placed and sealed. Ambronn (1), observing clinker grains, reported that fine crystal needles and little hexagonal platelets grew out from the borders of the grains. After some days he also observed much larger hexagonal plates, and gel formation. Keisermann (81) verified these observations and determined to his satisfaction the constituent oxides of the various formations, by means of staining methods which he developed. As previously mentioned (Section VII), he concluded that the fine needles and gel consisted of calcium silicate. He identified the large hexagonal plates as $Ca(OH)_2$, and the platelets as a calcium aluminate. He also observed and identified all these products, except the platelets, in hydrating calcium silicate "melts". In this latter case the needle crystals developed only after some days, and they radiated from seemingly random points throughout the slide. Keisermann also obtained the needles through reaction of finely powdered, dried $Ca(OH)_2$ and hydrous silica. After 4 weeks, both large and small needles appeared, as well as a gel mass. Blumenthal's microscopical work (16), also mentioned earlier, was largely a repetition of Keisermann's, but in the identification of the fine needles it was apparently more thorough. Blumenthal confirmed Keisermann's identifications.

The fine needles, described as calcium silicate, were apparently first seen and identified by Le Chatelier (see Section II). Numerous other investigators (11, 27, 56, 61, 108, 135, 142, 154, 179) besides Ambronn, Keisermann, and Blumenthal have also observed them, generally in the hydration of cement or clinker. This is of considerable interest, since Klein and Phillips (85) and later Lerch and Bogue (122) reported their inability to find them though they devoted considerable effort to the attempt, using methods of observation recommended by Ambronn and Keisermann. Lay (108), who said that the needle crystals grew into platelets, may perhaps have mistaken platelets on edge for needle crystals; Blumenthal remarked that platelets on edge appeared at first like short thick needles. Perhaps some of the "silicate" needles reported in observations on cements were not silicate at all, but represented hydrated calcium aluminate or calcium sulfoaluminate. However, from the various descriptions there seems little reason to doubt that a considerable number of investigators did see crystal needles very similar to those which Keisermann and Blumenthal identified as calcium silicate and considered to be monocalcium silicate hydrate.

The extreme thinness of the calcium silicate needles has been stressed by various investigators. Ambronn (1) said that the needles could be seen in ordinary observations only by the use of a very narrow cone of light but that they stood out plainly under dark-field illumination obtained by use of a Zeiss paraboloid condenser. Keisermann (81) also said that they were hard to find. Scheidler (154) said that if the clinker was cooled too rapidly the needles would not form. That their number is dependent on the proportion of water used is indicated by Keisermann (81), Muth (135), and others. Muth (135) reported that the needles formed most voluminously when the proportion of water was large, were stable in solution that was well below saturation with $Ca(OH)_2$, and hence could not be simply $Ca(OH)_2$. However, Tippmann, who also observed needles, at first (179) called them calcium silicate but later (180) contended that they were $Ca(OH)_2$ and that hydrous silica does not react chemically with lime at ordinary temperatures. This last conclusion was strongly disputed by Kühl (95), who said in the course of the controversy that he considered that Keisermann and Blumenthal had demonstrated that their needle crystals contained silica. Ward (185) observed occasional long thin needles during the hydration of $K_2O \cdot 23CaO \cdot 12SiO_2$ (169); he also observed vermiform outgrowths from hydrating $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ grains. He considered both formations to be dicalcium silicate hydrate. Brown (22) has observed birefringent fibrous fringes on hydrating particles of $3CaO \cdot SiO_2$ in portland cement clinker of relatively low $3CaO \cdot Al_2O_3$ content. Also, in one instance he has obtained distinct needle crystals radiating from hydrating clinker grains. No gel formation was visible when the needles were first observed. Later, when gel had developed, the needles appeared less perfect than at first. The clinker grains used in preparing the slide were of rather uniform size, about 10 to 15 microns in diameter.

It has been reported that the silicate needles later change into gel. Scheidler (154) apparently said this, and Kühl and Mann (98) believed that Muth did also. Muth (135), however, wrote simply of the needle formation being depressed, with greater development of gel, in preparations made up with less water than others. Indeed, Muth said that he observed no change in the needles, but he attributed to Ambronn a belief that they finally decomposed. This seems also to be a misunderstanding, for Ambronn (1) said that he could still see the needles through gel that formed around them. Pulfrich and Linck (142) reported that the needles were destroyed when the gel formation began. They considered it erroneous, however, to say that the needles "change" to gel, because transformation of gel to crystals is not reversible. Pulfrich and Linck appear, nevertheless, to have believed that the needles are composed of calcium silicate. They thought that needles are formed only when a large proportion of water is used. This agreed with their observations, in connection with which they used glycerol solutions in order to reduce the amount of water while still employing enough liquid for convenient observations. Glycerol, however, is not an appropriate diluent (98, 179), since it can dissolve lime and may also have a specific effect on crystal formation. Like Keisermann, Pulfrich and Linck also obtained needle crystals through reaction of hydrous silica and $Ca(OH)_2$, but only after adding solid $Ca(OH)_2$ and shaking at 100°C. The needles developed after some weeks, growing eventually to lengths of 0.1 to 0.2 mm.

Le Chatelier (119) remarked years after his original investigation that crystalline hydrated calcium silicate had never been observed microscopically. This statement was challenged by von Glasenapp (57) on the basis of his own observations of needle formations. It is doubtful, however, that Le Chatelier really intended to retract his original identification (117) of crystal needles as $CaO \cdot SiO_2 \cdot 2.5H_2O$. His later discussion relates to the tendency of initially formed colloidal particles to develop into visible crystals, a change to be expected on thermodynamic grounds but for which there is indeed no direct microscopic evidence in the case of the hydrous calcium silicate. Kühl (94) has noted an increase in the birefringence of hydrated portland cement over a period of twenty years, but there is no assurance that this resulted from the growth of silicate crystals. After hydrating calcium silicates for two years, Bogue and Lerch (18) could obtain no indications of crystalline hydrates, other than $Ca(OH)_2$, either microscopically or by x-ray diffraction.

Owing to the fineness of the needles, their optical properties are hard to determine. Probably for this reason, and no doubt also because of actual differences between some of the needles observed, the accounts of the optical properties have been conflicting. A number of investigators (1, 16, 61, 135, 142) have reported that the needles are somewhat birefringent, whereas others (27, 57) found them to be wholly isotropic. Tippmann (179) said that the needles which he observed were originally isotropic but became birefringent, and this he attributed to surface carbonation. Blumenthal (16) reported that needles observed by him were birefringent with negative elongation. Pulfrich and Linck (142) distinguished thin and thick needles. The refractive indices of the thin needles were $\alpha = 1.498$, $\gamma = 1.500$ (error = 0.0015). Those of the thicker needles were a few units higher in the third decimal place. Grün, Tremmel, and Kunze (61) reported long thin needles with refractive indices of 1.47 to 1.53. The needles that Ward obtained with $K_2O \cdot 23CaO \cdot 12SiO_2$ had a mean refractive index close to 1.62, showed parallel extinction, and were biaxial, optically positive.

The gel itself has some interesting features, as observed under the microscope. Ambronn (1) noted that the gel forms at first as minute globules, and he describes how these droplets form along the needles like strings of pearls, and on the platelets and unchanged clinker grains. Finally the drops unite to form hyaline coatings. Lay (108) reported that the droplets also form on the slide and cover glass, often in regular geometric arrangements. The observations of both of these authors were made on slides of hydrating cement clinker. Ward (185), examining hydrating $3CaO \cdot SiO_2$ grains, observed that the vermiform growths, previously mentioned, changed in a few hours to round, gel-like particles. Eventually the individual grains were surrounded either by individual globules or by a more or less continuous gelatinous coating of the same refractive index as the globules (1.43 to 1.49). Finally, the $3CaO \cdot SiO_2$ grains were completely converted into individual lumps of gel which, as seen in a photomicrograph, appear to have retained essentially the original grain shapes. In this connection Brown (22) reports that hydration products of the silicates in portland cement quite commonly appear as pseudomorphs of the original mineral grains, in agreement with the observations reported by Nacken (137). Ward (185) reported that a thin groundmass gel formed like an etch over the cover glass, and on the slide carrying the $3CaO \cdot SiO_2$. This gel had a refractive index of 1.54, about the same as for the principal gel formation that he found in thin sections of hydrated cement.

Brown and Carlson (23) found that air-dried gels produced by the hydration of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ had refractive indices of 1.515 and 1.53, respectively. Heating either gel for 1 hr. at 900°C. raised its index to 1.695. From observations on hydrating portland cement these authors noted that gel and calcium hydroxide crystals began to form almost immediately. The gel was found to form around every cement particle and in a few days the spaces between the particles were filled with either gel or calcium hydroxide, except for occasional air bubbles. The gel, in all stages of dryness, gave no evidence of fine crystallinity. In the moist form it was found to react instantly with carbon dioxide, forming minute calcium carbonate crystals. This carbonation was observed to occur more easily "in the gel than in areas of recognized crystalline hydroxide."

Most of the early microscopical studies of the hydration of calcium silicates were made with the object of learning something about the hydration of portland cement in construction work. However, since in order to observe the individual grains a much larger proportion of water was needed than is used in ordinary cement practice, the significance of the results for cement technology has been challenged, notably by Kühl (91). Kühl later (96) adopted the point of view that excess water could be used if it was saturated with $Ca(OH)_2$, like the solution in cement paste. He pointed out, however, that the latter solution generally contains an appreciable concentration of alkalies, as well as $Ca(OH)_2$. In early studies, Read (148) used limewater in making microscopic observations, but few investigators have followed this lead. So far as the present writer has been able to determine, the recorded occurrences of the silicate needles are all for studies in which plain water was used. Ward (185), who did not obtain silicate needles (except when hydrating $K_2O \cdot 23CaO \cdot 12SiO_2$), found that in saturated $Ca(OH)_2$ solution the reactions of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ were similar to those in plain water but that there was a noticeable decrease in the rate of hydration. The addition of gypsum, potassium hydroxide, or sodium hydroxide had no visible effect on the nature of the reaction products of the silicates. Nacken (137) noted no decrease in the rate of hydration of his $3CaO \cdot SiO_2$ when he used saturated limewater, but did when he also added solid $Ca(OH)_2$. As previously mentioned, Muth (135), who treated cement with various proportions of water, found the least needle formation in his least dilute preparations. In cement pastes, in which the proportion of water is still less, Kühl (94) contends that the hydration products are almost entirely ultramicroscopic in structure, except for occasional $Ca(OH)_2$ crystals.

Bessey (11) has reported an accidental preparation of weakly birefringent parallel or fibrous material having a refractive index of about 1.51 and a composition corresponding approximately to $3CaO \cdot 2SiO_2 \cdot aq$. This apparently crystalline material formed on sand–lime brick stored in water, apparently at room temperature.

The electron microscope has been used by Radczewski, Müller, and Eitel (144) to study the hydration of $3CaO \cdot SiO_2$. Some very fine crystal needles were observed which were believed to be a calcium silicate hydrate. Sliepcevich, Gildart, and Katz (159) took electron microscope pictures of products obtained by evaporating solutions in which particles of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ had been hydrating (separately). Besides $Ca(OH)_2$ particles, and some ill-defined material, rhombic formations were noted. These latter closely resembled calcium carbonate crystals obtained by carbonation of $Ca(OH)_2$, but the authors believed them to be of different composition, since carbon dioxide was thought to have been excluded. The $Ca(OH)_2$ produced by hydrolysis is shown by electron microscopy to form initially in large measure as rounded particles of nearly perfect circular (and half-circular) cross section, 1 micron and less in diameter (143, 144, 159).

In hydrothermal studies, at temperatures above 100°C., crystalline hydrated calcium silicates are readily formed, as established both by x-ray and by the microscope. Non-equilibrium forms are apparently easily obtained, with the result that the early studies in this field do not establish the equilibria. In a recent paper (71), however, a tentative phase diagram is advanced. It is not the purpose here to attempt a systematic account of hydrothermal studies, which were reviewed in 1938 by Thorvaldson (174) and by Büssem (25). It may be remarked, however, that hydrates having $CaO:SiO_2$ ratios of 1:1, 3:2, 2:1, and 3:1 have all been reported, the latter, however, having only been obtained out of contact with liquid water.

The only direct connection between these hydrothermal products and the products obtained at ordinary temperatures is that provided by Mlle. Foret's x-ray identifications (49) of her crystalline $CaO \cdot SiO_2 \cdot aq$ (obtained at about 130°C.) with Chassevent's (31) gelatinous calcium silicate.

Refractive indices of "artificially hydrated calcium silicates" are tabulated by Bessey (11). Values of 1.59 to 1.64 are reported for various preparations of $2CaO \cdot SiO_2 \cdot aq$. For $3CaO \cdot 2SiO_2 \cdot aq$ and $CaO \cdot SiO_2 \cdot aq$ the values range from 1.57 to 1.60. X-ray and optical data on natural and synthetic crystalline hydrated calcium silicates are given by McMurdie and Flint (125) and Flint, McMurdie, and Wells (45) (see also an earlier paper by Vigfusson (184)).

XV. BOUND WATER

Isotherms showing the water content of the gelatinous calcium silicate in relation to vapor pressure, at apparent equilibrium, have been obtained by a few investigators. The curves are shown in figure 3. Bessey's curve is for his 17°C. $1CaO:1SiO_2$ preparation (11). Lefol's preparations (120) had 1.2 to 1.3 moles of CaO per mole of SiO_2, but on the basis of Le Chatelier's work (117) the CaO in excess of $1CaO:1SiO_2$ was regarded as adsorbed. Berchem's tests (9) were made on $3CaO \cdot SiO_2$ that had hydrated for 71 days and must therefore have contained much free $Ca(OH)_2$. Powers (139) used specimens obtained by hydrating $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ for over three years, the original aqueous

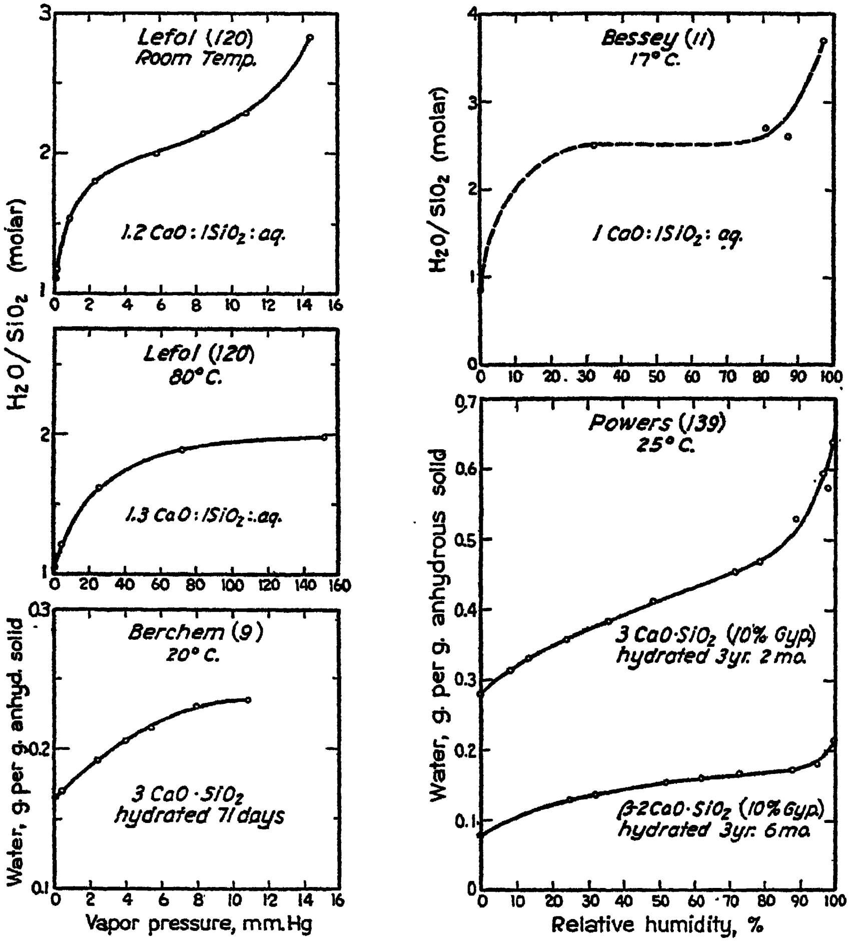

FIG. 3. Isothermal dehydration curves for hydrous calcium silicate preparations. For Lefol's room-temperature curve the three highest points were obtained by rehydration.

pastes having been made up with 10 per cent additions of gypsum which is believed to have remained primarily unreacted. Probably a large part of the β-$2CaO \cdot SiO_2$ (at the centers of the original particles) remained unhydrated. All curves of figure 3 were obtained by progressive *dehydration*, except that in Lefol's room-temperature curve H_2O/SiO_2 values above 2 were obtained by rehydration.

Lefol concluded from the flatness of the curves in the neighborhood of 2 moles of water, and their termination at about the 1 mole level, that definite hydrates having 1 and 2 moles of water are indicated. His initial preparations, containing about 2 moles of water, were obtained by drying *in vacuo* over calcium chloride. Lefol also made other tests in which the $1.2CaO:1SiO_2:2H_2O$ preparation was heated, with uniform increase in temperature, and was weighed at intervals without interrupting the treatment. The curve obtained showed that the loss of water began at about 80°C. and that the rate of loss decreased rather abruptly at about 160°C., when the water content was about 1.2 moles. Under continuous heating at 145°C. (in a separate test) the water was reduced to about 1 mole and then remained constant. Hence, 1 mole of water was assumed to be combined with 1 mole of $CaO \cdot SiO_2$. Since more than 1 mole of CaO was present per mole of SiO_2, it would appear that the extra lime, if adsorbed (as was assumed), would remain at 145°C. as $Ca(OH)_2$, and that correction should be made for the water thus combined in estimating the water in $CaO \cdot SiO_2 \cdot aq$. However, Lefol makes no mention of any such correction, and it is clear that the data in figure 3 are uncorrected.

Bessey, also, concluded from his curve that the gel probably contains a definite hydrate, evidently $CaO \cdot SiO_2 \cdot 2.5H_2O$, though this is not stated. Berchem, however, regarded his curve as determined by the pore structure of the gel, and believed it to be without significance as to the hydrates present. Powers drew similar conclusions, reporting that his curves represented water held not as definite hydrates, but in a manner typical of colloidal gels.

Isobars showing the retention or loss of water by hydration products of calcium silicates at a succession of temperatures are shown in figure 4. The $3CaO:1SiO_2$ preparation used by Krauss and Jörns (89) was evidently a mixture of $2CaO \cdot SiO_2$ and CaO, since it was quenched from a melt (37). Meyers (131) maintained a pressure of only about 0.1 micron of mercury; thus he obtained the "steps," L, in his curves at lower temperatures than did Krauss and Jörns, who maintained a pressure of 7 mm. of mercury. These steps can be attributed to dehydration of $Ca(OH)_2$, since they occur at about the temperatures to be expected, in view of the pressure maintained. Except for these steps and the irregularities up to about 100°C., the curves show the same colloidal, adsorptive, or zeolitic character as the isotherms. Forsén's (53) isobars resemble those obtained by Meyers, but they were not carried to temperatures high enough to dehydrate $Ca(OH)_2$ if any was present.

Isotherms and isobars have also been obtained for hydrated portland cement. In their indications regarding hydrous calcium silicates they are reasonably consistent with the curves obtained directly on such materials. Hence they will not be considered further here, except for the remark that the isotherms show hysteresis, the dehydration curve not being retraced throughout by the rehydration curve (140). As mentioned earlier, it is primarily the dehydration isotherms that have been determined for the hydrous calcium silicates.

Although, because of their rather flat central portions, isotherms such as Lefol's do appear to suggest definite hydrates, it should be realized that very similar curves are obtainable with a number of finely porous adsorptive substances, and in many such cases the curve is attributed wholly to adsorption and capillarity (24). Indeed, the isotherms show a sensitivity of water content to vapor pressure that is similar to the sensitivity of CaO/SiO_2 to CaO concentration. Thus, the bonding of the water remains uncertain in much the same way that the bonding of the CaO is uncertain. Even the fact that the gel still retains a considerable amount of water at seemingly negligible vapor pressure is not good assurance that this residual water is all held as some definite hydrate (109).

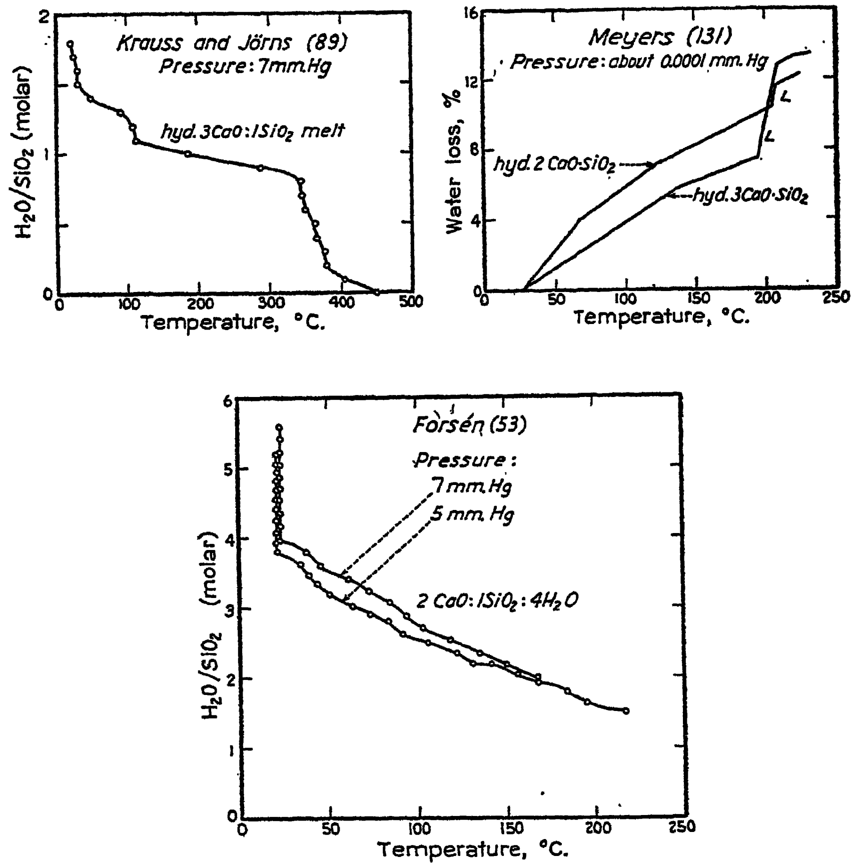

FIG. 4. Isobaric dehydration curves for hydrous calcium silicate preparations

Nevertheless, it is evident that a considerable amount of water is firmly held. The amounts that remain at very low vapor pressure, and at rather high temperatures, attest to this. Cirilli (33), like Lefol, found that on heating an approximate $1CaO:1SiO_2:aq$ preparation at a fairly high temperature (130°C.) 1 mole of water was retained even on prolonged heating. Roller and Ervin (152) cite these findings of Cirilli and Lefol as indications that 1 mole of water is constitutional.

Cirilli (33) also found that upon heating directly to a higher temperature (160°C.) actually more water was retained than at 130°C. and that the product had an unusually high heat of solution. He concluded, therefore, that something other than simple dehydration was involved. In his early studies (33) Cirilli obtained the same heat of solution, in acid, for $1CaO:1SiO_2:aq$ both before and after drying from 70 to 13.4 per cent (or 1 mole) of water. Removal of this last mole of water at 650°C. apparently gave SiO_2 and $2CaO \cdot SiO_2$, since the latter was indicated by the x-ray spectrum. Cirilli later (34) reported heats of solution for samples of $1CaO:1SiO_2:aq$ having water contents ranging from 1 to 26 moles. In contrast with the earlier results, the data show a gradual rise in the heat of solution with reduction in water from 26 to 2 moles and a much more pronounced rise thereafter.

Other investigators have determined the "fixed water" by some arbitrary standardized procedure. Work and Lasseter (191) compared methods of four different types, using hydrated cement as test material. These methods were: oven drying for 1–4 hr. at 90°, 105°, and 130°C.; grinding with absolute alcohol at room temperature; extraction at the boiling point with benzene (80–82°C.), toluene (111°C.), and xylene (139°C.); and evacuation at 4–10 mm. of mercury at room temperature for 2 hr. The cement had been mixed originally with 30 per cent of its weight of water, but the exact amount retained at the time of the tests is not stated. Results are reported as the losses of water during the various treatments, expressed in percentages of the weight of the sample after ignition at 850°C. Except for the evacuation at 4–10 mm., which gave a distinctly low result (9.2 per cent water), and the 90°C. oven drying, which gave about 13 per cent water, the various tests gave remarkably similar results: 14.6 to 18.0 per cent water loss. On the basis of these tests, oven drying for 3 hr. at 105°C. was adopted, and the water retained at 105°C. but lost at 850°C. was called the combined water. During drying, a stream of dry carbon dioxide-free air was slowly passed over the crucible. Tricalcium silicate, made up as a very wet mix and hydrated for six months, showed, per mole of SiO_2, 1.8 moles of water of hydration plus 1.4 moles in $Ca(OH)_2$ as determined by the glycerol–alcohol method (123).

Boque and Lerch (18) determined "fixed water" by drying for 24 hr. at 105°C. and igniting at 1000°C. Their $3CaO \cdot SiO_2$, ground till 90 per cent passed the 200-mesh sieve and made up with 50 per cent of its weight of water, showed at six months 1.1 moles of water of hydration besides 0.8 mole in $Ca(OH)_2$ (glycerol–alcohol method, after heating at about 530°C.). At two years the water in $Ca(OH)_2$ had not increased, but the water of hydration was 1.6 moles.

Hedin (67) took precipitated calcium silicates and dried them over calcium chloride in an atmosphere free from carbon dioxide, then over phosphorus pentoxide, and finally in a drying oven at 105°C., drying to constant weight each time. Results are given below:

$CaO:SiO_2$	MOLES WATER RETAINED ON DRYING		
	Over $CaCl_2$	Over P_2O_5	At 105°C.
1.07:1	2.48	1.72	1.33
1.50:1	3.48	2.73	2.01
2.07:1	4.01	3.56	2.64

Hedin's product that was extracted with ethylene glycol (see Section XII) was dried in a desiccator at 105°C. It then had the composition $1.5CaO:1SiO_2:1.57H_2O$.

Forsén (52) gives the following compositions for products "dried at 105°C.":

$1.06CaO:1SiO_2:1.17H_2O$
$1.13CaO:1SiO_2:1.07H_2O$
$2.06CaO:1SiO_2:4.27H_2O$
$1.99CaO:1SiO_2:3.79H_2O$
$2.05CaO:1SiO_2:3.84H_2O$

Cirilli (34) states that his products that had CaO/SiO_2 ratios equal to 1:1 and 3:2 retained water corresponding to $CaO \cdot SiO_2 \cdot H_2O$ and $3CaO \cdot 2SiO_2 \cdot 3H_2O$ when they were dried to constant weight *in vacuo* over phosphorus pentoxide and potassium hydroxide.

Nacken (137) does not describe his drying procedure but reports water values that vary erratically, as, for example: $1.96CaO:1SiO_2:2.1H_2O$ and $2.10CaO:1SiO_2:0.97H_2O$. In his various preparations of varied CaO/SiO_2 ratio the range in moles of water per mole of silica was 0.97–3.9. Keevil and Thorvaldson (80) state that the apparent monocalcium silicate obtained by hydrolysis of $3CaO \cdot SiO_2$ retains 1 mole of water after drying to constant weight over quicklime. Koyanagi's supposed monocalcium silicate retained approximately 2 moles of water after drying to constant weight in a vacuum desiccator.

XVI. THE PHASE DIAGRAM FOR THE SYSTEM $CaO–SiO_2–H_2O$

It will be evident from the foregoing sections that there is not as yet any possibility of drawing a precise ternary diagram for the system $CaO–SiO_2–H_2O$ that would be generally accepted. Indeed, if particle size and adsorptive capacity are as important as has been suggested, it does not appear that a single precise diagram that does not take these properties into account could represent adequately all the valid relationships when the solids remain as colloids. However, the results of individual experimenters can of course be represented in this way, if the data are sufficiently complete.

Flint and Wells (47), and also Nacken (137), made use of ordinary rectangular plots of SiO_2 concentration *versus* CaO concentration, with the compositions of equilibrium solids suitably indicated, and with the anhydrous solids of the system represented by lines drawn from the origins with slopes equal to the CaO/SiO_2 values. As Nacken had not actually determined SiO_2 concentrations, he drew his solubility curve along the CaO axis.

Bessey (11) employed the equilateral triangular plot to represent a combination of his own and other data. He treated Flint and Wells' relationships at the lower lime concentrations as metastable. Equilibrium solids were assumed to be $Ca(OH)_2$, $2CaO \cdot SiO_2 \cdot aq$, solid solutions between $3CaO \cdot 2SiO_2 \cdot aq$ and $CaO \cdot SiO_2 \cdot aq$, and $SiO_2 \cdot aq$.

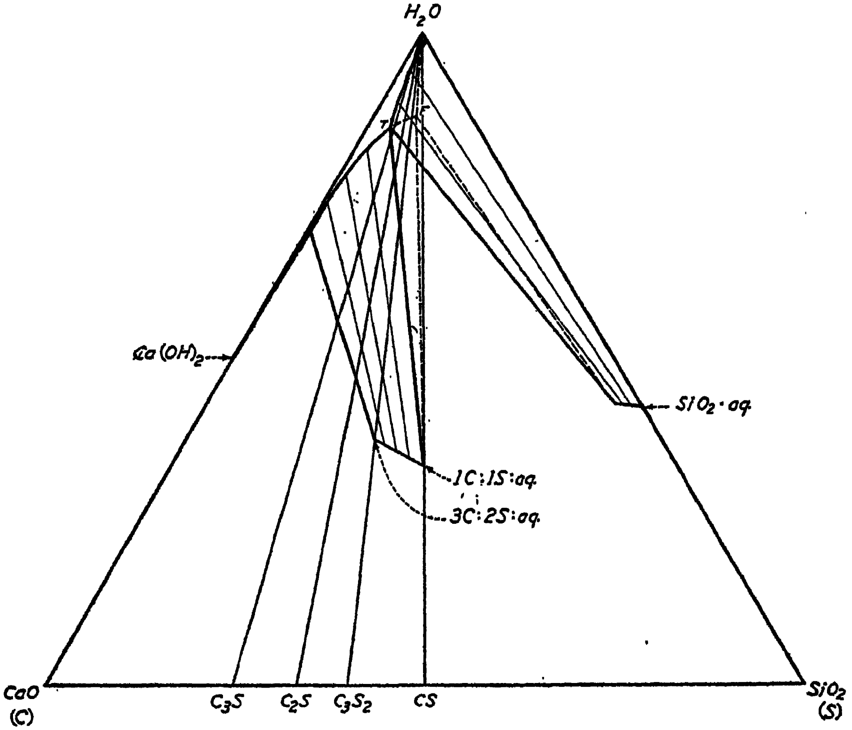

FIG. 5. Hypothetical phase diagram for the system $CaO-SiO_2-H_2O$ at ordinary temperature with exaggerated solubilities (presented only as an aid to discussion).

Owing to low solubilities of the solids involved, the equilateral triangle is not a very satisfactory means for presenting data for the $CaO-SiO_2-H_2O$ system if the components are expressed in uniform units. However, because it is so well known, it is used in figure 5 to illustrate a few points, though with solubilities exaggerated and the water contents of the solids arbitrarily assumed. The hydrous solids are taken as: $Ca(OH)_2$, a product varying in composition from $1CaO:1SiO_2:aq$ to $1.5CaO:1SiO_2:aq$, and $SiO_2 \cdot aq$ with some adsorbed $Ca(OH)_2$. These assumptions regarding solids are in substantial agreement with the apparent indications of Thorvaldson and Vigfusson's data (curve 13, figure 1).

Tie lies connect points on the solubility curves with the equilibrium solids. The appearance of the plot in this respect does not disclose whether the solids are solid solutions or single compounds with adsorbed lime; either assumption can be made if no further evidence is available. Lines drawn to the water apex from the points representing the anhydrous calcium silicates cut through the other tie lines wherever the indicated liquid–solid equilibria can be produced simply by adding water to the anhydrous compounds. For example, curves 13, figure 1, seem to indicate that addition of enough water to $3CaO \cdot SiO_2$ will give as the resultant solid phase a compound lower limed than $CaO \cdot SiO_2 \cdot aq$, perhaps $SiO_2 \cdot aq$ as assumed in figure 5; accordingly, the invariant point *T* is placed so that the lines from it to the two solids $CaO \cdot SiO_2 \cdot aq$ and $SiO_2 \cdot aq$ are both cut by the line from $3CaO \cdot SiO_2$ to H_2O. This is simply a graphical representation of the fact that if the equilibrium solids are to have CaO/SiO_2 lower than in $3CaO \cdot SiO_2$, the solution phase must have $CaO/SiO_2 > 3$. (The dashed lines of figure 5 are considered later.)

TABLE 7

Concentrations of solutions in equilibrium with calcium silicate gels as determined by Roller and Ervin (152)

CaO	SiO_2	CaO	SiO_2
millimoles/liter	*millimoles/liter*	*millimoles/liter*	*millimoles/liter*
0.99	0.690	2.81	0.107
1.01	0.660	2.84	0.112
1.18	0.442	2.92	0.108
1.52	0.293	3.16	0.0924
1.64	0.267	3.41	0.0821
1.83	0.215	3.52	0.0777
2.55	0.140	4.54	0.0510
2.69	0.119	6.24	0.0337
2.70	0.114	8.76	0.0238

If the deduction just stated could be accepted as valid, it would mean that Thorvaldson and Vigfusson had established an important fact regarding the CaO/SiO_2 value of the solution at the invariant point without having made a single determination of SiO_2 in solution. However, the deduction is in poor agreement with experimental data obtained by certain of the other investigators. It is, for example, at variance with many of the data of table 4. Though presence of silica sol accounts, no doubt, for a number of the high silica values that have been reported, it is believed that considerable reliance can be placed in the data of Flint and Wells (47) and Roller and Ervin (152) that were obtained at the high-lime side of the invariant point in question. The data of these authors seem to be in reasonable agreement in this region; the values tabulated by Roller and Ervin are given in table 7.[12]

[12] In figure 4 of Roller and Ervin's paper the Ca^{++} scale is displaced one division to the right, as compared with the tabulated values.

According to table 7, the CaO/SiO_2 ratio exceeds 3 only at CaO concentrations above about 1.2 millimoles per liter, or 0.07 g. CaO per liter. This concentration is above that at the invariant points indicated by some of the investigations, though below the CaO concentrations at the invariant points suggested by the curves (No. 13, figure 1) based on the data of Thorvaldson and Vigfusson. As a further step in analysis, it is advisable to consider the possible consequences of the fact that curves 13 of figure 1 are based on the assumption that the SiO_2 in solution can be neglected in computing CaO/SiO_2 for the solid. A trial has shown that even if the Thorvaldson and Vigfusson data are corrected for dissolved SiO_2, by use of table 7, the points at which the curves break are but

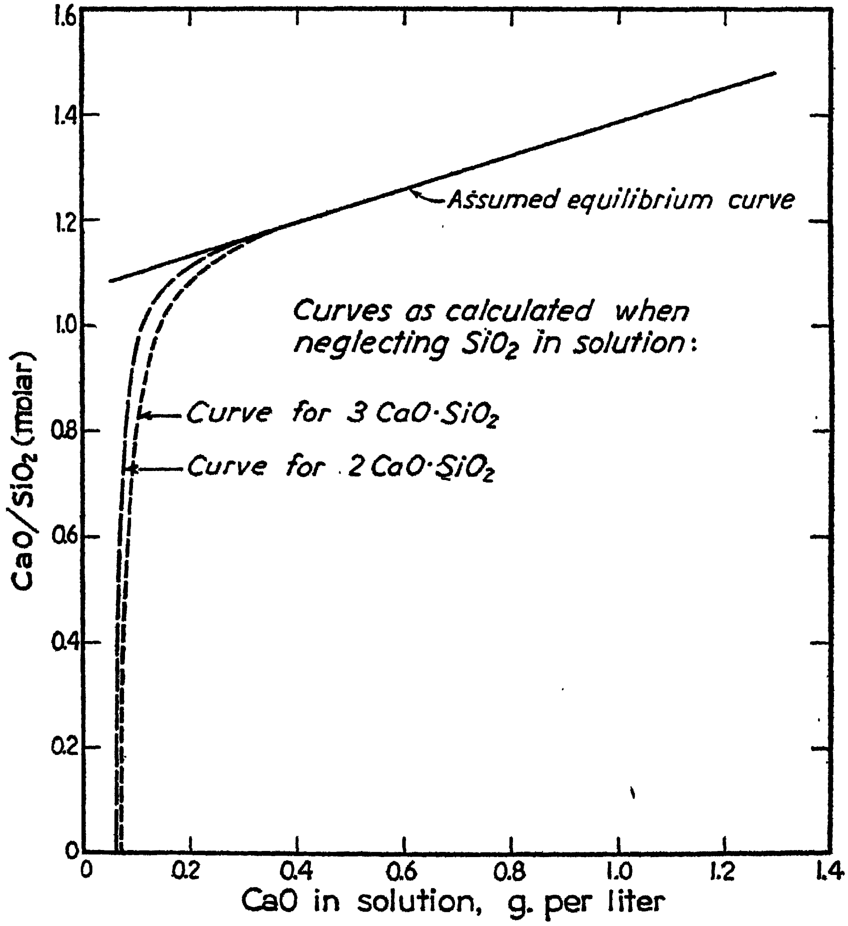

FIG. 6. Hypothetical curves for CaO/SiO_2 for solids *versus* CaO in solution. Used to illustrate the possible effects of neglecting SiO_2 in solution when each preparation is made simply by adding water to anhydrous calcium silicate.

little affected. However, this comparison is not very significant because the authors conceded poor analytical accuracy at the lower CaO concentrations. The accuracy would indeed have to be poor in order to account for the large discrepancy between the two curves in this region. Consequently, it is of interest to consider what effect the neglect of SiO_2 could have in an analogous hypothetical case.

In figure 6 the upper straight line represents a hypothetical relationship between CaO/SiO_2 in solids and CaO in solution, one that resembles the upper portion of Thorvaldson and Vigfusson's curve but is extended to lower CaO concentrations. The lower curves, one for $3CaO \cdot SiO_2$ and one for $2CaO \cdot SiO_2$, were developed from the straight curve by assuming SiO_2 in solution in accordance with the solubility data of Roller and Ervin, letting CaO/SiO_2 be the ratio of CaO in solids to *total* SiO_2 as in curves 13 of figure 1. It is evident that this method, applied at sufficiently low concentrations of CaO, produces inflections rather similar to those obtained by Thorvaldson and Vigfusson, though no invariant point has been assumed. Furthermore, though the curves for $3CaO \cdot SiO_2$ and $2CaO \cdot SiO_2$ are much closer together than curves 13 of figure 1, they are in the same order. This outcome does not prove that Thorvaldson and Vigfusson did not establish an invariant point, but it does suggest uncertainty in the matter. In this connection it is to be recalled (from Section II) that Michaelis reported that cements treated with large volumes of water seemed to hydrolyze only so far as to produce the monocalcium silicate. Leaching of course is a different matter and apparently leads eventually to $SiO_2 \cdot aq$ (122). Nacken (137) who, like Thorvaldson and Vigfusson, simply added water to $3CaO \cdot SiO_2$ also obtained one point for which CaO/SiO_2 is much less than 1 (see curve 16, figure 1), but he also did not actually analyze the solid, and did not correct for SiO_2 in solution.

Point *F*, with its dashed connecting lines, has been entered in figure 5 to show how the invariant point, represented by *T*, must be shifted in order to locate it in accordance with Bessey's ratio of CaO/SiO_2 for the invariant solution (11). The point must, in any case, fall at the high-lime side of the $CaO \cdot SiO_2$ —— H_2O line, since the monocalcium silicate hydrate evidently dissolves incongruently, giving a solution with higher CaO/SiO_2 values than the compound itself (11).

Figure 5 shows the chemical compositions of all the anhydrous calcium silicates that have regions of stability in the $CaO–SiO_2$ system at atmospheric pressure (37, 146). Not shown by the plot is the fact that $CaO \cdot SiO_2$ and $2CaO \cdot SiO_2$ can exist in two and three enantiotropic forms, respectively. The $3CaO \cdot SiO_2$ does not have a thermodynamically stable form at room temperature (29, 116, 130), but it is stable in a practical sense and shows no dissociation when maintained at room temperature (111). Tricalcium silicate reacts with water much more rapidly than the $2CaO \cdot SiO_2$ and the still lower-limed calcium silicates (1a, 6, 18, 47, 85, 113, 191). Also, the high-temperature (α- and β-) forms of $2CaO \cdot SiO_2$ are distinctly more reactive than the low-temperature γ-form. In some preparations, the γ-form and the lower-limed calcium silicates appear almost inert toward water. The β-$2CaO \cdot SiO_2$ has been cited earlier in this review as the form that is present in portland cement. There has, however, been some controversy on this point (112, 166), since the α- and β-forms have been reported as having similar properties (64, 146). Complex twinning has been thought to characterize the α-form, but Insley (70) has suggested that crystals that show this intricate twinning may simply be β-$2CaO \cdot SiO_2$ that was originally of the α-form, the complex twinning having been caused by the inversion. Greene (60) has recently presented evidence that supports this view.

None of the anhydrous calcium silicates are stable in contact with their aqueous solutions at room temperature (47, 113). Nor do any of the hydrated calcium silicates that may be formed dissolve congruently in water, since leaching of even the highest-limed calcium silicate, $3CaO \cdot SiO_2$, leaves finally a solid phase consisting only of hydrous silica (122). However, Flint and Wells (47) found that $CaO \cdot SiO_2$ (pseudowollastonite?) forms a supersaturated solution of great stability. They report a time study of the compositions of the solutions that form when the various calcium silicates are shaken with water.

Hydrous silica has been represented in figure 5 simply by $SiO_2 \cdot aq$. The existence of definite hydrates of silica has been reported from time to time, only to be denied by other investigators. For example, Thiessen and Koerner (172) and Spychalski (161) reported vapor-pressure isotherms which suggested the presence of definite hydrates in specially prepared silica gels, but their results could not be confirmed by Weiser, Milligan, and Coppoc (189). As discussed by Hurd (69), several investigators have concluded that within the elemental solid units of a silica gel the silica is probably anhydrous. However, whether the water is held entirely superficially or not, a considerable amount, namely, 0.3 to 1.0 mole per mole of SiO_2, is firmly held, apparently as hydroxyl groups (30, 41). It has been pointed out that this is a rather large amount to be held entirely superficially (41).

XVII. EFFECTS OF ALKALIES AS ADDITIONS TO THE $CaO-SiO_2-H_2O$ SYSTEM

The possible effects of the addition of alkali oxides to the system $CaO-SiO_2-H_2O$ are of considerable interest to cement technologists. Portland cements contain amounts of chemically combined $Na_2O + K_2O$ that range from 0.2 per cent or less to well in excess of 1 per cent. A large part dissolves quickly; indeed, as much as half of the potash and 10 to 20 per cent of the soda have been extracted from certain cements of normal fineness in only 7 min. simply by making a paste with water (79). It is reasonable to expect that the alkalies in the liquid phase in concrete may sometimes amount to 10 g. Na_2O, or K_2O, per liter (78). To a considerable extent these alkalies form the corresponding hydroxides in solution (115), thus raising the pH above that for saturated $Ca(OH)_2$. The initial basicity is limited, however, by reaction between alkali hydroxides and the gypsum in cement (151). Also, if the clinker contains sulfur trioxide, much of the potassium oxide may enter the solution as potassium sulfate (170). Within about a day, however, the sulfate is largely removed from solution through calcium sulfoaluminate formation (115), and then practically all dissolved alkali is present as hydroxide.

Shaw and MacIntire (158) reported a few experiments on the effects of sodium chloride and sodium hydroxide as additions to the $CaO-SiO_2-H_2O$ system. They found that both of these alkali compounds decreased the CaO/SiO_2 value of the gel and that the sodium hydroxide was particularly effective. A few additional observations were reported at the symposium held by cement chemists and technologists in Stockholm in 1938 (141). Lea (110) remarked that the incongruent solubility of hydrated calcium monosilicate is higher in 3 per cent sodium chloride solution than in water. Thorvaldson (176) stated that at a given alkalinity the CaO/SiO_2 value of the gel is lower when sodium sulfate is present. Bessey (14) speculated that alkali hydroxide in solution might lower the CaO/SiO_2 value of the equilibrium solid. To test this, Forsén (53) introduced 0.5 g. of his supposed $2CaO \cdot SiO_2 \cdot aq$ into the filtrate from a suspension of $3CaO \cdot SiO_2$ in 2 per cent potassium hydroxide solution. After 72 hr. he found the $2CaO \cdot SiO_2 \cdot aq$ to be unchanged. Forsén (52) also reported that hydrolysis of $3CaO \cdot SiO_2$ proceeds rapidly even in potassium hydroxide "of considerable concentration." Later, Hedin (67) reported that the rate of this reaction is almost independent of the pH of the solution. His strongest potassium hydroxide solution was 10 per cent, or about 1.8 normal. In contrast, Lerch and Bogue (122) considered that they effectively stopped the hydrolysis of $3CaO \cdot SiO_2$ and of lower-limed silicates by using sodium hydroxide solution of pH 13.72 (evidently about 1.1 normal) at 30°C.

Kalousek (78) has now reported an extensive systematic study of selected portions of the system $Na_2O-CaO-SiO_2-H_2O$ at 25°C. He determined the compositions of solutions and gels that coexist with solid $Ca(OH)_2$. Also, at selected sodium oxide concentrations, he investigated preparations that precipitate a gel, but not any $Ca(OH)_2$. The sodium oxide concentrations ranged from 0.2 to 152 g. per liter. The initial materials were solutions of sodium silicate and sodium hydroxide, crystalline calcium hydroxide, and additional water, as needed. Preparations were contained in ceresin-lined flasks which were shaken a number of times during the first day and once daily thereafter. Analyses were made after 2 to 18 weeks. These periods were selected arbitrarily, but "neither the solid phase nor solution underwent any determinable changes in time periods of from several weeks to several months." Precipitates were washed with aqueous alcohol solutions of progressively increased alcohol content, up to 95 per cent alcohol, and finally with ether, after which the residue was left in the laboratory air for 1 day. The alcohol contents of the wash solutions were chosen by testing the original filtrate and selecting the maximum alcohol content which just failed to produce a suspension. The solutions and most of the precipitates were analyzed for Na_2O, CaO, and SiO_2. Combined water was determined by difference. The gross CaO/SiO_2 values, as determined for the solids, were checked against values computed from the mix proportions and the analyses of the filtrates.

These gross CaO/SiO_2 values as established in single tests did not disclose the CaO/SiO_2 values for the gels when solid $Ca(OH)_2$ was present. Accordingly, the CaO/SiO_2 values for the gels in equilibrium with solid $Ca(OH)_2$ were determined by making in each instance a series of preparations in which CaO was varied while maintaining a fixed concentration of Na_2O. The plots of CaO/SiO_2 for solids *versus* CaO (or SiO_2) in solution showed, by the breaks in the curves, the values of CaO/SiO_2 that were sought.

When the Na_2O in solution was 0.2 g. per liter, the gel in equilibrium with solid $Ca(OH)_2$ had the composition $0.003Na_2O:1.94CaO:1.0SiO_2:3.2H_2O$. As the Na_2O concentration of the solution was raised while still maintaining solid $Ca(OH)_2$, the gel composition altered, with Na_2O increasing and CaO decreasing, until at 20 g. Na_2O per liter the gel composition was approximately $0.25Na_2O:1.0CaO:1.0SiO_2:2.8H_2O$. Except for minor and rather random deviations, this composition remained practically constant as the concentration of Na_2O was increased to 101 g. per liter. (At higher Na_2O concentrations CaO/SiO_2 for the gel was not determined.) With increase in the Na_2O in solution, the SiO_2 concentration increased to over 1.1 g. per liter. In other tests in which the solutions were not saturated with $Ca(OH)_2$ the SiO_2 in solution was increased much more. The gels in equilibrium with these solutions had CaO/SiO_2 values well below 1, but their Na_2O/SiO_2 values held close to 0.2 over wide ranges in Na_2O and SiO_2 in solution.

The author discusses the significance which his results may have with respect to concretes. He considers that the alkali content of the cement keeps the CaO/SiO_2 value for the gel considerably lower than it would otherwise be. He finds, however, no agreement between his results and the compositions that have been reported for the gelatinous accumulations found in concretes made with destructive "reactive" aggregates. Analyses of these latter gels show mainly silica, alkali oxides, and water. Kalousek attributes their compositions to the semi-permeable nature of hardened cement paste and the resultant osmotic effects, as assumed by Hansen (65).

Kalousek does not comment upon the possible modes of retention of the Na_2O and CaO in the gels. It may be noted, however, that the results obtained in the presence of solid $Ca(OH)_2$ show considerable agreement with the theory of adsorption on $CaO \cdot SiO_2 \cdot aq$, extra CaO in the gel being replaceable with Na_2O, though not on an equimolal basis. The CaO/SiO_2 value of 1.94:1 that was obtained when Na_2O was only 0.2 g. per liter is of interest for comparison with results of other studies, as given in table 5. For zero Na_2O concentration the value should evidently be about the same, though a little higher.

XVIII. CALCIUM SILICATE GEL IN HYDRATED PORTLAND CEMENT

The siliceous product of portland cement hydration is generally believed to be a hydrous lime-silica gel such as is formed in the $CaO-SiO_2-H_2O$ system at high $Ca(OH)_2$ concentration (113). As previously mentioned, this viewpoint has inspired much of the work that has been done on the $CaO-SiO_2-H_2O$ system. In the present review, it has led to the inclusion of some experiments and observations made directly on hydrating cements. Therefore, a brief discussion of some of the evidence for calcium silicate gel in hydrated cement is thought advisable.

It is significant, first of all, that the calcium silicates, $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$, modified a little by some solid solution, generally constitute 75 per cent, or more, of normal portland cement clinker. Moreover, various investigations (for example, (18)) have shown that these compounds are themselves hydraulic cements reacting with water to form products comparable in eventual strength development to the hydrated portland cement. To be considered also are the staining tests which have indicated in microscopic studies (16, 81, 137) that the cement gel contains CaO and combined SiO_2, but no Al_2O_3, and further that the Al_2O_3 forms a calcium aluminate hydrate. There is also other good evidence that the Al_2O_3 forms non-siliceous products in cement: calcium sulfoaluminate and calcium aluminate hydrate (113). Oxides other than CaO, SiO_2, and Al_2O_3 amount to less than 10 per cent of the clinker and often none of them exceeds 3 per cent; it seems reasonable therefore to assume in hydrated portland cement a hydrous calcium silicate gel, one that is modified, however, by some content of alkalies, as indicated by the work of Kalousek (78), and by intimate intermixture with other products of cement hydration (23). Indeed, rather decisive evidence as to the constitution of the gel in cement has now been provided by Strätling (165), who found that the gelatinous hydration product of cement, when separated from unreacted material, shows the same x-ray spectrum as hydrous calcium silicate gel. Büssem (25) reports that Brandenberger (19) obtained similar evidence regarding portland cement, but Brandenberger's account shows that hydrated $3CaO \cdot SiO_2$, rather than cement, was examined.

It may be noted, next, that some investigators have presented evidence that has led them to suppose that hydrous calcium aluminosilicates may be formed during the hydration of cement. In that case, the amount of calcium silicate gel would be somewhat reduced.

Strätling (165), working with zur Strassen, found that, after some months, products that were formed by shaking $3CaO \cdot SiO_2$ and $3CaO \cdot Al_2O_3$ (two constituents of cement) together in water showed not only the x-ray spectrum that he obtained with hydrous calcium silicate but also another. This latter spectrum was one that he had also obtained under other circumstances, circumstances which led him to believe that it was a quaternary phase, $2CaO \cdot Al_2O_3 \cdot SiO_2 \cdot aq$, though he did not obtain it in a pure state. However, though he obtained the x-ray spectrum of the hydrous calcium silicate in a test on hydrated portland cement, he did not then obtain the spectrum for the assumed quaternary phase. He believed, nevertheless, that at equilibrium the alumina should be represented in $2CaO \cdot Al_2O_3 \cdot SiO_2 \cdot aq$.

Flint, McMurdie, and Wells (46) discovered the garnet–hydrogarnet series of solid solutions. These are compositions of the general formula $3CaO \cdot (Al_2O_3, Fe_2O_3) \cdot 3(2H_2O, SiO_2)$ that are formed, under suitable conditions, at elevated temperature (about 100°C. and higher). The authors speculated as to whether similar compounds might be formed in hydrated portland cement at ordinary temperature but stated, explicitly, that they had no proof of such occurrence.

Later, Flint and Wells (48) discovered hydrous calcium silicoaluminates analogous to the calcium sulfoaluminates which occur in one or more forms in hydrated portland cement. One of these silicoaluminates was produced from certain binary mixtures of cement compounds that had been placed in water and stored for three years at room temperature. However, no success was had with mixtures of $3CaO \cdot SiO_2$ and $3CaO \cdot Al_2O_3$. The authors suggest that the silicoaluminates may possibly play a rôle in cement hydration, perhaps through the formation of solid solution with the sulfoaluminates. In this connection they mention a calcium sulfosilicoaluminate reported by Lafuma (104). Inspection of Lafuma's paper shows however, that the author did not establish that he had only a single solid phase. The composition he reported was $9CaO \cdot 4SiO_2 \cdot Al_2O_3 \cdot 7CaSO_4 \cdot 80H_2O$.

It must be concluded that there is still no direct evidence that a hydrous calcium aluminosilicate is formed during the normal hydration of portland cement.

XIX. SUMMARY OF RESULTS AND CONCLUSIONS FROM PAST INVESTIGATIONS

Results of the various investigations of the system $CaO-SiO_2-H_2O$ at ordinary temperature, and the conclusions drawn from them, are summarized below:

1. A common method of studying phase relations in the system $CaO-SiO_2-H_2O$ is to plot the molar CaO/SiO_2 value for the solid phase against the CaO concentration of the solution. The curves of this kind for data obtained at ordinary temperatures in different investigations have varied widely, as shown by figure 1. In general, however, the curve rises rapidly before the CaO concentration reaches 0.15 g. per liter, and it attains a CaO/SiO_2 value of about 0.85–1.1. The rise then becomes much more gradual but continues, throughout, while the CaO concentration is further increased. The values of molar CaO/SiO_2 attained in the solids, in various investigations, when the solutions were saturated with $Ca(OH)_2$ (about 1.2 g. CaO per liter) have ranged from 1.4 to 2.

2. The solid products are ordinarily colloidal and give no evidence of crystallinity when observed through the light microscope. An x-ray spectrum of a few lines has been obtained, however, by several different investigators, and the spectrum is reported to be the same, within the limits of error, for products of varied molar CaO/SiO_2 (above 1).

3. Though some of the individual differences in the curves may be attributable to differences in the initial materials, essentially the same general shape of curve as described above has been obtained in instances where the hydrous lime-silica products were formed by reaction between silica sol or gel and limewater, sodium silicate and limewater, and anhydrous calcium silicate and water.

4. Divergent results can no doubt be accounted for partially, and perhaps largely, on the basis of incomplete attainment of colloidal equilibrium, and other insufficiencies in experimental work. The possible effects of moderate differences in temperature are also to be considered. However, since the solid products are colloidal, it is to be recognized that only colloidal equilibrium is attained in any case. Differences in adsorptivity and in solubility, as a result of differences in particle size, may therefore be important.

5. Apparently, the initial abrupt rise of the curve denotes chemical combina-

tion of the silica with lime. This inference is supported by various lines of evidence, including the fact that the silica in the $1CaO:1SiO_2$ product is completely acid-soluble. Whether more than one hydrous calcium silicate is formed during this initial rise is in dispute, since the rise is found to be practically vertical in some investigations but not in others. Moreover, values reported for the SiO_2 concentrations of the solutions in this region have shown pronounced variation both within the individual investigations and with respect to different investigations. The variation has been considered important by some investigators, while by others it has been attributed simply to the presence of colloidal silica.

6. Many investigators believe that the initial abrupt rise in the CaO/SiO_2 value for the solid represents formation of $CaO \cdot SiO_2 \cdot aq$. This judgment now has the support of x-ray evidence, since crystalline $CaO \cdot SiO_2 \cdot aq$ prepared hydrothermally at 100–140°C. has been reported (49) to give the same spectrum as the gel obtained at room temperature. Formation of still lower-limed calcium silicate hydrates has also been assumed.

7. Whether CaO held in excess of 1 mole per mole of SiO_2 is chemically combined with the silica or is simply adsorbed is a disputed matter. Thermal data have been advanced in support of chemical combination, but applications of the solubility product principle and the Freundlich adsorption isotherm have favored adsorption. In any case, the fact that the additional lime is acquired over a wide range in CaO concentrations in solution appears to point to either adsorption or some kind of solid solution.

8. The SiO_2 in solution increases, at first, as the CaO concentration is increased from zero, and is indicated to attain its maximum value in the region of abrupt rise in the CaO/SiO_2 ratio of the solids. Thereafter, it falls rapidly and is indicated to be less than 0.01 g. SiO_2 per liter at 30°C., when the CaO concentration is 0.15 g. CaO per liter. Maximum values as high as 0.3 g. SiO_2 per liter have been reported, but whether they are equilibrium values has been questioned. Metastability and silica sol formation have been suggested. Some investigators believe that the maximum SiO_2 concentration is only about 0.05 g. SiO_2 per liter or less. Owing to the strong decrease in SiO_2 concentration on the high CaO side of the point of maximum concentration, nearly all the CaO is present as $Ca(OH)_2$ in solutions in which CaO is more than about 0.2 g. per liter.

9. None of the anhydrous calcium silicates are stable in contact with their aqueous solutions at room temperature. Also, none of the hydrated calcium silicates that may be formed in such solutions appear to be capable of dissolving congruently in water, for leaching of even the highest-limed calcium silicate, $3CaO \cdot SiO_2$, leaves finally a solid phase consisting only of hydrous silica.

10. By reason of hydrolysis, a large part of the curve for CaO/SiO_2 in solids *versus* CaO in solution can be established simply by adding water (in varied proportions) to either of the high-limed calcium silicates, $3CaO \cdot SiO_2$ or $2CaO \cdot SiO_2$, none of the soluble products being removed from the system. Indeed, it has been indicated that either compound, employed in this way, can advance the hydrolysis until the solid phase is lower limed than $CaO \cdot SiO_2 \cdot aq$

and is, perhaps, $SiO_2 \cdot aq$. If this is possible, it means that the molar CaO/SiO_2 in solution must exceed 3:1 at the invariant point where $CaO \cdot SiO_2 \cdot aq$ and the lower-limed compound are in equilibrium. Such a ratio at this point is, however, larger than is indicated by some of the other evidence, and it appears that the matter needs further investigation.

11. The gelatinous solid having a CaO/SiO_2 value of less than 2, that has been obtained in apparent equilibrium with saturated $Ca(OH)_2$ solution in many investigations, indicates that both $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$ should react with water with precipitation of $Ca(OH)_2$, a fact that has also been established by direct experiment. In spite of this, β-$2CaO \cdot SiO_2$ made up with water as a paste appears to hydrate with very little further hydrolysis after saturating the solution with $Ca(OH)_2$. Also, much of the data for hydration of $3CaO \cdot SiO_2$ in a paste indicates that only a third, or less, of the CaO is liberated as $Ca(OH)_2$. It thus appears that during hydration in pastes the hydrolysis is often less than for equilibrium. Here again, however, the data are not wholly consistent. Also, the methods of determining $Ca(OH)_2$ are apparently somewhat faulty.

12. At least at concentrations of CaO above that at which CaO/SiO_2 for the solid reaches 1, all the SiO_2 in the liquid phase is apparently in true solution. Probably at the higher CaO concentrations the silicate ions are primarily those of orthosilicic acid, H_4SiO_4, though this is not fully established. At lower CaO concentrations some investigators believe that the silicate ions associate to form dimers. At the higher CaO concentrations the molar ratio of "silicate CaO" to SiO_2 has been reported to exceed 1 in solutions supersaturated with the calcium silicate.

13. The amount of water held by the hydrous calcium silicates under varied circumstances has been less investigated than the relative amounts of CaO and SiO_2. However, the water content is found to decrease with each decrease in external vapor pressure, after the manner of many other colloidal materials. In compositions approaching $CaO \cdot SiO_2 \cdot aq$ the first molecule of water appears to be rather firmly held and may be constitutional.

14. Under some circumstances, mainly during the hydration of the anhydrous silicates in the presence of considerable water, fine needle crystals have formed, which, mainly on the basis of staining tests, have been called calcium silicate, or silicate hydrate, and have frequently been represented as $CaO \cdot SiO_2 \cdot aq$. The conditions under which these needles form have not been well established and various investigators have been unable to obtain them.

15. The addition of sodium hydroxide to the $CaO–SiO_2–H_2O$ system when the solution is maintained saturated with $Ca(OH)_2$ has been found to reduce the CaO-to-SiO_2 ratio of the gelatinous product that is in equilibrium with the solid $Ca(OH)_2$. The CaO in the gel in excess of 1 mole per mole of SiO_2 is found to be replaceable by about 0.25 mole of Na_2O.

16. Portland cement, which contains about 75 per cent of $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$, produces on hydration a gelatinous product which is evidently largely a lime–silica gel such as is formed in the system $CaO–SiO_2–H_2O$, except that it presumably contains some alkali also. However, all the products of

cement hydration are closely intermingled. The solution phase is saturated with $Ca(OH)_2$ and is commonly supersaturated during the early stage of the hydration.

XX. NEED FOR FURTHER RESEARCH

It is evident from the foregoing that in spite of the large amount of work done, some of it apparently skillful and painstaking, few precise numerical data on compositional relationships in the system $CaO-SiO_2-H_2O$ at ordinary temperature can yet be advanced as necessarily having general validity. This does not mean, however, that the various researches have been of little value. Indeed, the broad outlines of the picture appear to have been rather well established. Also, as previously discussed, it is probably inherent in the colloidal nature of the solids formed in this system that the gross compositions must vary somewhat with circumstances other than the final temperatures, vapor pressures, and solution concentrations. Nevertheless, it should eventually be possible to advance the knowledge of the $CaO-SiO_2-H_2O$ system at ordinary temperature far beyond its present state.

Some of the past work could advantageously be repeated with modifications suggested by the uncertainties that now exist. Baylis' observations (7) led him to conclude that the concentration of the initial silica gel affects the amount of CaO that can be taken up. If further systematic study confirms this, it would seem to limit the possibilities with respect to the nature of the final product. Hedin (67) believes that colorimetric determination of SiO_2 gives only the silica in true solution, but Krasil'nikov and Kiselev (88) appear to be less confident. Some method of distinguishing adequately between colloidal and dissolved silica could aid materially in establishing more thoroughly the invariant points of the system. Thorvaldson and Vigfusson (178) believe that their anhydrous silicates, $3CaO \cdot SiO_2$ and β-$2CaO \cdot SiO_2$, hydrolyzed in large volumes of water until the molar CaO/SiO_2 of the solid phase was less than 1, but the composition of the final solid was obtained indirectly, and without allowance for SiO_2 in solution. As the result is of much interest, repetition of some of the experiments with direct analysis of final solids and solutions for both CaO and SiO_2 seems advisable. Flint and Wells (47) concluded that the relatively high-limed solutions of the $CaO-SiO_2-H_2O$ system have molar ratios of "silicate CaO" to SiO_2 greater than 1, but they worked with strongly supersaturated solutions. Perhaps development of analytical methods of sufficiently high precision may eventually enable this conclusion to be checked with solutions that are not supersaturated. Various investigators believe that they have observed needle crystals of calcium silicate hydrate which have formed at room temperature. Apparently, there is a reasonable possibility that by special techniques crystals of microscopic size could be formed at room temperature in quantities sufficient to enable their compositions and properties to be determined. Indeed, some of the crystalline calcium silicate hydrates obtained hydrothermally might well be further investigated with respect to equilibria at room temperature.

The past work has indicated that attainment of colloidal equilibrium is sometimes slow and uncertain. In future studies, the more promising of the methods already used can be selected and applied, with approach to equilibrium from both the high- and low-lime sides. More attention can be focused, also, upon the nature of the bond between the CaO and SiO_2 in the hydrous solids. Greatly needed, in this connection, is a reliable quantitative test for $Ca(OH)_2$ in the presence of the hydrous calcium silicate gel but the reliability of such a test is not easy to prove. Possibly one of the current methods is sufficiently accurate, but if so this needs to be better established. These methods have considerable utility, however, especially for relative purposes. It is, therefore, somewhat surprising that they have been little used (for $Ca(OH)_2$, as distinct from CaO) except in studies on hydrated pastes, and on commercial cements. Free $Ca(OH)_2$ tests applied in studies like those represented in figure 1 might provide information of distinct value. In view of the possibilities of adsorption it would be of interest to apply such tests to solids obtained at all $Ca(OH)_2$ concentrations.

The cause of the increase in the CaO/SiO_2 value of the solid with increase in the CaO concentration of the solution should also become more evident if electrokinetic tests were made and if more work were done with other metallic ions besides Ca^{++}. Ion-exchange studies would bear directly on the questions of solid structure and the state of combination of the CaO.

The calorimetric approach, employed by Cirilli, appears to warrant further use. Also, electron microscopy and both x-ray and electron diffraction can undoubtedly be used to further advantage. Radioactive calcium might also prove serviceable in studying adsorption.

In many of the investigations of the past the experimenters have contented themselves with determinations of the gross compositions of the colloidal products, in relation to solution compositions. In the future it should be possible to bring improved methods to bear upon the problem of the structure of the colloidal products and to distinguish more clearly between the chemical and physical forces that are involved.

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THE QUANTITATIVE ASPECT OF DIFFUSION IN ELECTROLYTE SOLUTIONS

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I. INTRODUCTION

There are few domains of physical science in which so much experimental effort over nearly a century has yielded so little accurate data as the field of diffusion in liquid systems. The numerical difficulties inherent in the computation of the diffusion coefficients from rate measurements, the elimination of turbulent flow, the very accurate control of temperature, and the analytical accuracy required are all contributing obstacles to the attainment of high accuracy.

In general, the theory of diffusion in liquid systems is undeveloped. It is difficult and complicated and, owing to the scarcity of experimental results of fundamental nature, has not made great progress. However, electrolytes possess certain properties which have made possible the development of an exact theory for very dilute solutions. As the concentrations of strong electrolytes approach zero, the properties of the ions and in particular the mobilities approach additivity. This behavior was utilized by Nernst (71) to compute the limiting value of the diffusion coefficient from the knowledge of the ionic mobilities at infinite dilution. On the other hand, the diffusion coefficient varies with the electrolyte concentration. In dilute solutions, this variation may be theoretically computed because of the fortunate fact that the character of the force between ions is known. The theory of concentrated solutions is obscure because the short-range forces of interaction between the ions and the forces between the ions and the solvent are such that the net influences of all these factors cannot be predicted by Coulombic interaction alone.

In recent years, the diffusion of molecules in liquid media has been the subject of numerous reviews both in journals and in books (13, 19, 30, 72, 107). The modern theory of diffusion of electrolytes and an analysis of the most accurate determinations of diffusion coefficients up to 1943 has been presented in a systematic but specialized form by Harned and Owen (41). More recently, at a colloquium under the auspices of the New York Academy of Sciences on "The Diffusion of Electrolytes and Macromolecules in Solution" (1), a number of specialists have considered many of the aspects of the subject. All of these contributions are devoted to restricted phases of the field and none includes the fundamental theory of irreversible processes.

In this review, as indicated by the preceding table of contents, we shall strive to coördinate into a logical and reasonably complete treatment the scattered contributions to the theory. Since diffusion comes under the same class of flow phenomena as electrical conductance and heat flow, we shall begin with the general theory of irreversible processes as a necessary foundation for the discussion of electrolytic conductance and diffusion (77, 78). Then, after the introduction of some necessary relations from the interionic attraction theory (17, 18), we shall develop the present theory of the diffusion of electrolytes (79, 80) and finally derive the "dissipation function" for electrolytic conductance and diffusion. This procedure will bring together for the first time the somewhat scattered literature of this subject and provide a phenomenological and theoretical background for the experimental part of the review.

Of the many experimental methods, we shall describe in some detail only those which have led to good numerical results. Fortunately we shall be able to establish values for the differential diffusion coefficient of potassium chloride in dilute aqueous solution (39, 40) which can be used in establishing the theory at concentrations from 0.001 to 0.01 normal and also as a standard for the calibration of diaphragm cells. It is important to note that this makes possible a new field of investigation, since no previous measurements of differential diffusion coefficients of high accuracy have been reported below 0.05 normal concentration.

We conclude with some general considerations concerning the extension of theory to systems of more than two components and with a summary of the present status of the quantitative aspect of the subject.

II. Fundamental Theory and Definitions[1]

1. *Fick's first law*

Although diffusion in liquid solutions was observed by Parrot (81) as early as 1815, the first measurements of consequence were made by Graham (34, 35) and the first theoretical contribution by Adolph Fick (23) in 1855. If the quantity of solute which diffuses through unit area in unit time is $\mathbf{J}$, hereafter denoted "the flow," then Fick's first law takes the form

$$\mathbf{J} \equiv n\mathbf{v} = -\mathfrak{D}\nabla n = -\mathfrak{M}\nabla\mu \tag{1}$$

where n is the number of molecules of diffusing component per cubic centimeter, $\mathbf{v}$ its velocity, $\mathfrak{D}$ the diffusion coefficient of the solute, and μ its chemical potential. ∇ is the operator "del", and consequently ∇n and $\nabla\mu$ are the gradients of the concentration and chemical potential, respectively.

The flow is usually defined relative to a fixed plane of reference. This is sufficient for many applications, but is not valid for the general theory of the diffusion of electrolytes, since volume changes result from the mixing of the components. This difficulty may be remedied by defining the flow of a component relative to a frame of reference moving with the solvent (80). A bulk velocity defined by

$$\mathbf{v} \equiv \Sigma \mathbf{J}_i \bar{V}_i = \Sigma n_i \mathbf{v}_i \bar{V}_i \tag{2}$$

---

[1] As far as possible, the symbols used are those of Harned and Owen (41).

where $\bar{V}_0$, $\bar{V}_1$, $\cdots$ are the partial volumes of the components, may be employed in the general treatment of this phenomenon. For the present, we shall consider only the simpler case where the partial volume effect is omitted.

The form of Fick's law as stated by the last expression on the right of equation 1, which recognizes that the "force" causing diffusion is the gradient of the chemical potential of the diffusing substance, was first suggested by Hartley (42) and is a necessary basis for all our subsequent derivations.[2] Obviously,

$$\mathfrak{D} = \left(\frac{\partial \mu}{\partial n}\right)_{P,T} \mathfrak{M} \tag{3}$$

If the diffusion is restricted to motion in the x-direction only, then Fick's first law assumes the familiar form:

$$\mathbf{J} = -\mathfrak{D} \frac{\partial n}{\partial x} \tag{4}$$

This fundamental law for the flow of matter has the same form as that of the flow of electricity (Ohm's law) or the flow of heat (Fourier's law). Here the gradient of the concentration is analogous to the electromotive force and to the gradient of the temperature. In the discussion of the theory of simultaneous irreversible processes, we shall make full use of this analogy and refer to the gradients of concentration, electrical potential, and temperature as "forces" in a generalized sense.

We shall now introduce the relation which defines the activity coefficient ($\zeta \equiv a/n$):

$$\mu = \mu^0 + kT \ln \zeta n \tag{5}$$

where k is the gas constant per molecule, μ is the chemical potential of the diffusing substance, μ^0 is its chemical potential in a standard state, and n is the number of molecules per cubic centimeter. For a real solution, it follows from equations 3 and 5 that

$$n\mathfrak{D} = kT\mathfrak{M}\left[1 + n\frac{\partial \ln \zeta}{\partial n}\right] \tag{6}$$

In the limit when n approaches zero and ζ unity

$$\mathfrak{D} = \frac{kT}{n}\mathfrak{M} \tag{7}$$

The units in which the above equations have been expressed were chosen because they are convenient for theoretical developments.[3] To convert to more familiar units, we let

$$n = N\bar{n}; \qquad \mu N = \bar{\mu} \tag{8}$$

---

[2] Nernst's derivations assumed that the "force" which causes migration is the gradient of the osmotic pressure. Later, Schreiner (87) substituted a function of the activity for osmotic pressure.

[3] We draw attention to this rather confusing situation at this juncture so that the reader will have it in mind when more complicated derivations are involved.

where N is Avogadro's number and $\bar{n}$ is in mols per cubic centimeter. With these changes, equation 1 becomes

$$\bar{\mathbf{J}} = \bar{n}\mathbf{v} = -\overline{\mathfrak{M}}\nabla\bar{\mu} = -\left(\frac{\mathfrak{M}}{N^2}\right)\nabla\mu \tag{9}$$

where $\bar{\mathbf{J}}$ is the flow in mols per cubic centimeter per second. Further,

$$\mathfrak{D} = \frac{\partial\bar{\mu}}{\partial\bar{n}}\overline{\mathfrak{M}} = 1000\overline{\mathfrak{M}}\frac{\partial\bar{\mu}}{\partial c} \tag{10}$$

where c is the concentration in mols per liter. For an electrolyte dissociating into ν ions, equation 5 becomes

$$\mu = \mu^0 + \nu kT \ln \zeta_{\pm} n_{\pm} \tag{11}$$

and

$$c\frac{\partial\bar{\mu}}{\partial c} = \nu RT\left(1 + c\frac{\partial \ln y_{\pm}}{\partial c}\right) \tag{12}$$

where $y_{\pm}$ is the conventional mean activity coefficient defined on the molar scale (41). In these units the diffusion coefficient is related to $\overline{\mathfrak{M}}$ according to the equation:

$$\mathfrak{D} = \nu 1000RT\frac{\overline{\mathfrak{M}}}{c}\left(1 + c\frac{\partial \ln y_{\pm}}{\partial c}\right) \tag{13}$$

2. Fick's second law

By substituting the value of $n\mathbf{v}$ given by equation 1 in the equation of continuity

$$\frac{\partial n}{\partial t} + \nabla\cdot(n\mathbf{v}) = 0 \tag{14}$$

where $\nabla\cdot$ is recognized as the divergence, we obtain

$$\frac{\partial n}{\partial t} = \nabla\cdot\mathfrak{D}\nabla n \tag{15}$$

in general and for unidirectional flow

$$\frac{\partial n}{\partial t} = \frac{\partial}{\partial x}\mathfrak{D}\frac{\partial n}{\partial x} \tag{16}$$

This is the law for the differential diffusion coefficient as defined by equation 1. If $\mathfrak{D}$ is independent of the concentration, then

$$\frac{\partial n}{\partial t} = \mathfrak{D}\frac{\partial^2 n}{\partial x^2} \tag{17}$$

the familiar expression for Fick's second law.

3. Differential and integral diffusion coefficients

In general $\mathfrak{D}$ is a function of the concentration, so that it is necessary to distinguish between differential ($\mathfrak{D}$) and integral ($\mathfrak{D}_i$) diffusion coefficients.

The integral coefficient is an average value over a concentration range and is given by the integral

$$\mathfrak{D}_i = \frac{1}{n_2 - n_1} \int_{n_1}^{n_2} \mathfrak{D} \, dn \tag{18}$$

Of these two quantities, the differential diffusion coefficient is the one required by theory. Unfortunately, most measurements in the past yielded integral coefficients and true differential coefficients were rarely determined. Gordon (32) has aptly described the situation: "The resulting confusion has had some unfortunate consequences, since the 'diffusion constant' is usually only a constant in the Pickwickian sense of the term. Thus, one finds the quantity determined in an experiment in which a solution, initially decinormal, is allowed to diffuse into water, referred to as the 'diffusion constant for 0.1 *N*.' The fact that it could, with equal justice, be spoken of as the value at infinite dilution is conveniently ignored." We shall have occasion later to discuss Gordon's contribution to the clarification of this situation.

If measurements of sufficient accuracy can be obtained in the later stages of diffusion when $(n_2 - n_1)$ becomes small, the value of $\mathfrak{D}_i$ will approach closely the value of $\mathfrak{D}$ (58). Later, we shall discuss conductance measurements which take full advantage of this fact.

III. General Theory of Irreversible Processes (77, 78)

A detailed presentation of the theory of diffusion of electrolytes requires an exposition of the fundamental principles which underlie the behavior of combined irreversible processes. As illustrations of phenomena in which two or more irreversible processes may proceed simultaneously, we cite thermoelectric phenomena (Peltier effect), heat conduction in anisotropic media, and transport processes in electrolytic solutions. In the Peltier effect, the flow of electric current in a system of conductors of different metals will produce heat, and conversely, if the metallic junctions are maintained at different temperatures, electricity will flow. In anisotropic media, temperature gradients and the corresponding "flows" of heat will be different along the three space coördinates required to express the phenomena. General equations for this problem will involve combined relations of the flows in three directions. In electrolytic diffusion where two or more ions are present, interactions between the individual transport of the ions occur. The general theory of the reciprocal relations in such combined processes has been developed by Onsager (77, 78). It forms an essential background for the theory of electrolytic conductance and diffusion. A general but abbreviated presentation of this theory will be sufficient for our immediate purpose.

A. RECIPROCAL RELATIONS IN IRREVERSIBLE PROCESSES (THEORY OF ONSAGER)

1. Introduction to phenomenological relations

The earliest reciprocal relation of the coupled processes involved in thermoelectric phenomena (Peltier effect) was recognized by Thomson (Lord Kelvin)

(95). In the thermoelectric circuit let $\mathbf{J}_1$ be the electric current and $\mathbf{J}_2$ the heat flow. Let $\mathbf{X}_1$ equal the electromotive force which drives the electric current and $\mathbf{X}_2$ the "force" in corresponding units which causes the flow of heat and will be proportional to the gradient of the temperature. Then

$$\mathbf{X}_2 = -\frac{1}{T}\nabla T \tag{19}$$

where T is the absolute temperature. If the flows of current and heat happen to be independent, then

$$\begin{aligned} \mathbf{X}_1 &= R_1 \mathbf{J}_1 \\ \mathbf{X}_2 &= R_2 \mathbf{J}_2 \end{aligned} \tag{20}$$

where R_1 and R_2 are the electrical and thermal resistances, respectively. However, since the thermal and electrical processes may interact, it seemed probable to Thomson that the more complicated phenomenological equations

$$\begin{aligned} \mathbf{X}_1 &= R_{11}\mathbf{J}_1 + R_{12}\mathbf{J}_2 \\ \mathbf{X}_2 &= R_{21}\mathbf{J}_1 + R_{22}\mathbf{J}_2 \end{aligned} \tag{21}$$

must be employed. He also thought it plausible that the assumption of the reciprocal relation

$$R_{21} = R_{12} \tag{22}$$

was correct. This assumption has been generally accepted, since it has been confirmed by the best available measurements.

Thermoelectric phenomena in electrolytes investigated by Eastman (21) are examples of a situation where three simultaneous processes are superimposed on thermodynamic equilibrium. Imagine two chemically identical reversible electrodes connected by a solution of a suitable electrolyte and let these be immersed in thermostats at two different temperatures. An electromotive force will be generated and electric current $\mathbf{J}_1$ will flow. Simultaneously, diffusion of matter and transfer of heat will occur. If $\mathbf{J}_2$ is the flow of matter and $\mathbf{J}_3$ that of heat, the extension of the previous result for coupled processes will lead to the three phenomenological equations

$$\begin{aligned} \mathbf{J}_1 &= L_{11}\mathbf{X}_1 + L_{12}\mathbf{X}_2 + L_{13}\mathbf{X}_3 \\ \mathbf{J}_2 &= L_{21}\mathbf{X}_1 + L_{22}\mathbf{X}_2 + L_{23}\mathbf{X}_3 \\ \mathbf{J}_3 &= L_{31}\mathbf{X}_1 + L_{32}\mathbf{X}_2 + L_{33}\mathbf{X}_3 \end{aligned} \tag{23}$$

and reciprocal relations of the type

$$L_{12} = L_{21}; \qquad L_{13} = L_{31}; \qquad L_{23} = L_{32} \tag{24}$$

may be expected to occur.

Diffusion of s kinds of ions in a liquid medium is another phenomenon which may be expected to follow a similar pattern, provided interactions take place between ions of different kinds. In this case, equation 23 may be written in the general form

$$\mathbf{J}_i = -\sum_{k=1}^{s} \mathfrak{M}_{ik}\nabla\mu_k \tag{25}$$

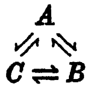
where $\nabla\mu_k$, the gradient of the chemical potential of each ion, is the analogue of the force $\mathbf{X}_k$, and $\mathfrak{M}_{ik}$ corresponds to L_{ik}. Here again, we suspect that

$$\mathfrak{M}_{ik} = \mathfrak{M}_{ki} \tag{26}$$

All these considerations suggest that there may be a general class of reciprocal relations in irreversible processes which may be derived from fundamental considerations of molecular mechanics.

2. Reciprocal relations in chemical reactions

Consideration of a triangular monomolecular chemical reaction in which three forms A, B, C of a substance exist simultaneously in the same phase yields analogous results. In this case, represented by the scheme



each of the forms may change into either of the others. According to the simple mass action law, the fraction of A molecules which changes into B in the short time Δt will be $k_{BA}\Delta t$, where k_{BA} is the proportionality constant. The rates of change of the concentrations n_A, n_B, n_C will be given by

$$\begin{aligned}
\frac{dn_A}{dt} &= -(k_{BA} + k_{CA})n_A + k_{AB}n_B + k_{AC}\bar{n}_C \\
\frac{dn_B}{dt} &= k_{BA}n_A - (k_{AB} + k_{CB})n_B + k_{BC}n_C \\
\frac{dn_C}{dt} &= k_{CA}n_A + k_{CB}n_B - (k_{AC} + k_{BC})n_C
\end{aligned} \tag{27}$$

If all the coefficients are $\gtrless 0$, then at equilibrium finite concentrations $\bar{n}_A$, $\bar{n}_B$, $\bar{n}_C$ are assured and these are given by relations of the type

$$\frac{d\bar{n}_A}{dt} = -(k_{BA} + k_{CA})\bar{n}_A + k_{AB}\bar{n}_B + k_{AC}\bar{n}_C = 0 \tag{28}$$

and

$$\bar{n}_A + \bar{n}_B + \bar{n}_C = n_A + n_B + n_C = n \tag{29}$$

In addition to these restrictions, physical chemists have assumed that when equilibrium is reached each individual reaction must be balanced microscopically. Thus, every transition from A to B will occur as frequently as the reverse transition from B to A. If then the equilibrium concentrations $\bar{n}_A$, $\bar{n}_B$, and $\bar{n}_C$ are known, three further relations

$$\begin{aligned}
k_{BA}\bar{n}_A &= k_{AB}\bar{n}_B \\
k_{CB}\bar{n}_B &= k_{BC}\bar{n}_C \\
k_{AC}\bar{n}_C &= k_{CA}\bar{n}_A
\end{aligned} \tag{30}$$

are imposed. These correspond to the reciprocal relations in the phenomenological equation (equation 23) and lead to one relation between the velocity constants:

$$k_{AC}k_{CB}k_{BA} = k_{AB}k_{BC}k_{CA} \tag{31}$$

Now this relation is not necessary to meet the thermodynamic requirement which will be satisfied when equilibrium is established by any set of positive values of the velocity constants. Indeed, other mechanisms can be postulated in which the detailed balancing of all the reactions does not occur. But these other mechanisms are not in accord with the idea that molecular mechanics and the dynamics of ordinary mechanics of conservative systems are fundamentally similar. With the exception of cases which involve Coriolis forces and external magnetic fields, the law of dynamic reversibility requires that if all the velocities of all the particles of a conservative dynamical system are reversed, the particles will retrace their former paths, thus reversing the previous succession of configurations. Applied to molecular mechanics, the assumption of dynamic reversibility implies that, if we wait until equilibrium is established, then every type of motion has a probability equal to that of its reverse. This implies the assumption made above that when the molecule A changes a given number of times to molecule B, the reverse transition of B to A takes place just as often.

In order to complete the analogy between the reciprocal relations (relations 24) and the phenomenological equations (equations 23), we must find the expressions analogous to $\mathbf{J}_1$, $\mathbf{J}_2$, $\mathbf{J}_3$ and $\mathbf{X}_1$, $\mathbf{X}_2$, $\mathbf{X}_3$. To this end we employ the thermodynamic equation for the free energy at constant temperature and pressure of a single reaction in a perfect gas system, namely:

$$F_{P,T} = F^0_{P,T} + RT\left[n_A \ln\frac{n_A}{\bar{n}_A} + n_B \ln\frac{n_B}{\bar{n}_B} + n_C \ln\frac{n_C}{\bar{n}_C}\right] \tag{32}$$

According to the condition specified by equation 29, it follows that:

$$\delta F_{P,T,n} = RT\left[\left(\ln\frac{n_A}{\bar{n}_A}\right)\delta n_A + \left(\ln\frac{n_B}{\bar{n}_B}\right)\delta n_B + \left(\ln\frac{n_C}{\bar{n}_C}\right)\delta n_C\right] \tag{33}$$

Let

$$x_A = n_A - \bar{n}_A, \text{ etc.} \tag{34}$$

and write δF in the form

$$\delta F = -\mathbf{X}_A\delta x_A - \mathbf{X}_B\delta x_B - \mathbf{X}_C\delta x_C \tag{35}$$

To obtain proportionality between the "forces" $\mathbf{X}_A$, $\mathbf{X}_B$, and $\mathbf{X}_C$ and the displacements x_A, x_B, and x_C, it is necessary to impose the restriction that the system is nearly in equilibrium. Thus

$$x_A << \bar{n}_A \text{ etc.} \tag{36}$$

in which case

$$
\begin{aligned}
\mathrm{X}_A &= -RT\ln\frac{n_A}{\bar{n}_A} \sim -\frac{RT}{\bar{n}_A}x_A \\
\mathrm{X}_B &\sim -\frac{RT}{\bar{n}_B}x_B \\
\mathrm{X}_C &\sim -\frac{RT}{\bar{n}_C}x_C
\end{aligned}
\tag{37}
$$

From equations 28 and 36, it follows that

$$\dot{x}_A = \frac{\mathrm{d}x_A}{\mathrm{d}t} = -(k_{BA} + k_{CA})x_A + k_{AB}x_B + k_{AC}x_C \tag{38}$$

or finally by application of equations 37:

$$
\begin{aligned}
\dot{x}_A &= (k_{BA} + k_{CA})\frac{\bar{n}_A}{RT}\mathrm{X}_A - \frac{k_{AB}\bar{n}_B}{RT}\mathrm{X}_B - \frac{k_{AC}\bar{n}_C}{RT}\mathrm{X}_C \\
\dot{x}_B &= -\frac{k_{BA}\bar{n}_A}{RT}\mathrm{X}_A + (k_{AB} + k_{CB})\frac{\bar{n}_B}{RT}\mathrm{X}_B - \frac{k_{BC}\bar{n}_C}{RT}\mathrm{X}_C \\
\dot{x}_C &= -\frac{k_{CA}\bar{n}_A}{RT}\mathrm{X}_A - \frac{k_{CB}\bar{n}_B}{RT}\mathrm{X}_B + (k_{AC} + k_{BC})\frac{\bar{n}_C}{RT}\mathrm{X}_C
\end{aligned}
\tag{39}
$$

If this be compared to equation 23, it is obvious that the analogy is complete, including all the reciprocal relations given by equations 24 and 30. Here $\dot{x}_A$, $\dot{x}_B$, $\dot{x}_C$ and the coefficients of X_A, X_B, X_C correspond to J_1, J_2, J_3 and L_{11}, L_{12}, $\cdots L_{ik} \cdots L_{33}$, respectively.

This conclusion suggests forcibly that general reciprocal relations may be deduced upon the basis of the principle of microscopic reversibility. To this end it is necessary to develop the theory by a method which involves no special mechanism. The only way in which this result can be achieved is by considering the *fluctuations* in a system which is in a normal state of thermodynamic equilibrium.

B. THE DERIVATION OF RECIPROCAL RELATIONS IN IRREVERSIBLE PROCESSES FROM THE GENERAL THEORY OF FLUCTUATIONS[4]

1. Formulation of the problem

By utilizing the fundamental equation of Boltzmann, relating entropy S and probability W

$$S = k\log W + \text{constant} \tag{40}$$

and by considering fluctuations in a set of variables $\alpha_1, \alpha_2, \cdots \alpha_n$ which measure displacements of matter, heat, and electricity, reciprocal relations in the simultaneous transport of these quantities will be derived by means of the principle of microscopic reversibility as applied to these fluctuations.

---

[4] We have not included in this review Onsager's derivation of reciprocal relations for the conduction of heat in anisotropic (triclinic) crystals.

We shall first define the quantities which correspond to the "flow" **J** and the "force" **X** and the coefficients in the phenomenological equation (equation 23). We start with the fundamental thermodynamic equation

$$\delta S = \frac{1}{T}(\delta E - \delta A) - \frac{\mu}{T}\delta m \tag{41}$$

where E, A, and m are the energy, work content, and mass of substance, respectively, and μ is the Gibbs chemical potential. The entropy will be determined by the heat displacements, α_1, α_2, α_3, so that

$$S = \sigma(\alpha_1, \alpha_2, \alpha_3) \tag{42}$$

Now, if there exists a uniform gradient of $1/T$ in the r-direction, a quantity of heat $\delta Q = (\delta E - \delta A)$ will be transferred a distance Δx_r

$$\delta S = \delta Q \cdot \Delta\left(\frac{1}{T}\right) = \delta Q \cdot \Delta x_r \frac{\partial(1/T)}{\partial x_r} = \delta\alpha_r \frac{\partial(1/T)}{\partial x_r} \tag{43}$$

and

$$\frac{\partial\sigma}{\partial x_r} = \frac{\partial(1/T)}{\partial x_r} \tag{44}$$

Similarly, if α is the displacement of matter

$$\frac{\partial\sigma}{\partial\alpha} = -\frac{\partial(\mu/T)}{\partial x} \tag{45}$$

and, if α is a displacement of electrical charge and **X** is the electrical field intensity:

$$\frac{\partial\sigma}{\partial\alpha} = \frac{\mathbf{X}}{T} \tag{46}$$

We note that the right members of these equations all correspond to the "forces" in equation 23 or equal $\mathbf{X}_1$, $\mathbf{X}_2$, $\mathbf{X}_3$, divided by T.

Now, the flows of heat, matter, and electricity are proportional to the gradients of the corresponding potentials. Thus,

$$\mathbf{J} \sim -\text{grad}\, T = -\nabla T \tag{47}$$

$$\mathbf{J} \sim \mathbf{X} \tag{48}$$

$$\mathbf{J} \sim -\nabla\mu \tag{49}$$

for the transport of heat, electricity, and matter, respectively. In general, these relations are given by

$$\frac{d\alpha_r}{dt} = \dot{\alpha}_r \sim \frac{\partial S}{\partial\alpha_r} \tag{50}$$

where $\dot{\alpha}_r$ is the flow J except for a volume factor. When different transport processes (n in number) interfere with each other, the flow is a function of the different kinds of displacements ($\alpha_1, \alpha_2, \cdots \alpha_n$) and

$$\dot{\alpha}_r = G_{r1}\frac{\partial\sigma}{\partial\alpha_1} + \cdots + G_{rn}\frac{\partial\sigma}{\partial\alpha_n} \quad (r = 1, 2, \cdots n) \tag{51}$$

and $S = \sigma(\alpha_1, \alpha_2, \cdots \alpha_n)$. We shall proceed to show that general reciprocal relations

$$G_{rs} = G_{sr} \tag{52}$$

can be deduced by application of the principle of microscopic reversibility.

2. General theory of fluctuations

Statistical mechanics applied to molecules provides an explanation of thermodynamic equilibrium as a statistical equilibrium of elementary processes among large numbers of molecules. The fundamental relation of Boltzmann

$$S = k \log W + \text{constant} \tag{53}$$

relating the entropy S with the thermodynamic probability W, is the most suitable as a basis for a theory of fluctuations from the state of equilibrium. W is the number of ways in which a given thermodynamic state may be realized and for a state of thermodynamic equilibrium is in general a very large number. Thus, a thermodynamic equilibrium state, specified by the energy and external variables (pressure, volume, etc.), is incompletely determined by molecular theory.

The Boltzmann constant k is 1.380×10^{-16} ergs per degree, so the probability of a deviation involving a change in entropy ΔS,

$$e^{\Delta S/k} \tag{54}$$

is appreciable when ΔS and k are of the same order of magnitude. Fluctuations of this order have been observed in a few favorable cases, as with liquids near the critical point and the Brownian movement of particles in liquids. To compute W, a complete molecular theory of the system is required. If the molecules of the system obey the laws of classical mechanics, W equals a gross extension in phase-space.

Consider a system of constant energy in which the external parameters (volume, pressure, etc.) and the number of atoms and molecules are fixed. If this system is isolated for a sufficient length of time to reach equilibrium, then we expect that it will have passed through all the states $\Gamma^1, \Gamma^2, \cdots \Gamma^l$ consistent with the conditions of isolation. Over a long time period t the system will spend a time t_r in the state Γ^r, and we may expect that $t_1, t_2, \cdots t_l$ will be proportional to the phase regions, $W_1, W_2, \cdots W_l$, respectively. Upon this assumption $W_1, W_2, \cdots$ are defined as proportional to $t_1/t, t_2/t \cdots$. Boltzmann's principle then may be stated in the form

$$S_r = k \log (t_r/t) + \text{constant} \tag{55}$$

where S_r is the entropy of the state Γ^r and t_r/t is the probability of this state.

It is necessary to define the probability in terms of the measurable variables $\alpha_1, \alpha_2, \cdots \alpha_n$. For this purpose, we introduce the distribution function $f(\alpha_1, \alpha_2, \cdots \alpha_n)$ and to meet the statistical requirements the probability of the state Γ^r will be equal to the integral of this function over the region

$$\begin{aligned} &\alpha_1^{(r)} < \alpha_1 < \alpha_1^{(r)} + \Delta\alpha_1 \\ &\cdot \\ &\cdot \\ &\cdot \\ &\alpha_n^{(r)} < \alpha_n < \alpha_n^{(r)} + \Delta\alpha_r \end{aligned}$$

Equation 55 then becomes

$$S_r = k \log f(\alpha_1^{(r)}, \alpha_2^{(r)}, \cdots \alpha_n^{(r)}) + k \log \Delta\alpha_1, \Delta\alpha_2, \cdots \Delta\alpha_n + \text{constant} \quad (56)$$

The magnitudes $\Delta\alpha$ should be of the same order as the fluctuations, in which case the product, $\Delta\alpha_1\ \Delta\alpha_2 \cdots \Delta\alpha_n$, varies so little for a change of entropy $S' - S''$ that the quantity $k \log [(\Delta\alpha_1' \cdots \Delta\alpha_n')/(\Delta\alpha_1'' \cdots \Delta\alpha_n'')]$ is negligible in comparison to $k \log [f(\alpha_1', \alpha_2', \cdots \alpha_n')/f(\alpha_1'', \alpha_2'', \cdots \alpha_n'')]$. Consequently, equation 56 may be written

$$S_r = k \log f(\alpha_1^{(r)}, \alpha_2^{(r)}, \cdots \alpha_n^{(r)}) + \text{constant} \quad (57)$$

Up to the present, an actual thermodynamic state is defined completely by the variables $\alpha_1, \alpha_2, \cdots \alpha_n$. Einstein (22) pointed out that the Boltzmann relation is valid for cases when this definition is not complete and may be applied to fluctuations from this equilibrium state. The probabilities of the different states will be of different magnitudes. Of these, a chosen thermodynamic state given by a set of variables, $\alpha_1, \alpha_2, \cdots \alpha_n$, will occur much more frequently than all the other states. Thus, we adopt the convention that the greatest entropy occurs when the variables have the values $\alpha_1', \alpha_2', \cdots \alpha_n'$ and this we denote by

$$\sigma_{1\cdots n}(\alpha_1', \alpha_2', \cdots \alpha_n') \quad (58)$$

The corresponding thermodynamic state will be specified by

$$\Gamma_{1\cdots n}(\alpha_1', \alpha_2', \cdots \alpha_n') \quad (59)$$

Then

$$\sigma_{1\cdots n}(\alpha_1', \alpha_2', \cdots \alpha_n') = k \log f(\alpha_1', \alpha_2', \cdots \alpha_n') + \text{constant} \quad (60)$$

measures the probability for finding the variables $\alpha_1, \cdots \alpha_n$ with the values $\alpha_1', \cdots \alpha_n'$.

For a single variable α_p, we may write

$$\sigma_p(\alpha_p) = k \log f_p(\alpha_p) + \text{constant} \quad (61)$$

where $\sigma(\alpha_p')$ is the greatest possible entropy when $\alpha_p = \alpha_p'$. Upon differentiation of this relation, we obtain

$$k \frac{df_p(\alpha_p)}{d\alpha_p} = f_p(\alpha_p) \frac{d\sigma_p}{d\alpha_p} \quad (62)$$

Further

$$\int_{-\infty}^{\infty} f_p(\alpha_p)\,\mathrm{d}\alpha_p = 1 \tag{63}$$

Now let the entropy $\sigma_p(\alpha_p)$ be a maximum when α_p equals α_p^0 and assume that $(\alpha_p - \alpha_p^0)f_p(\alpha_p)$ approaches zero for large values of $|\alpha_p - \alpha_p^0|$. If this latter assumption is not true, the function must have an infinite number of maxima. With these premises, an important result is obtained by evaluating the average

$$\begin{aligned}\overline{(\alpha_p - \alpha_p^0)\,\mathrm{d}\sigma_p/\mathrm{d}\alpha_p} &= \int_{-\infty}^{\infty} (\alpha_p - \alpha_p^0)(\mathrm{d}\sigma_p/\mathrm{d}\alpha_p)f(\alpha_p)\,\mathrm{d}\alpha_p \\ &= k\int_{-\infty}^{\infty} (\alpha_p - \alpha_p^0)(\mathrm{d}f(\alpha_p)/\mathrm{d}\alpha_p)\,\mathrm{d}\alpha_p \\ &= k[(\alpha_p - \alpha_p^0)f_p(\alpha_p)]_{-\infty}^{\infty} - k\int_{-\infty}^{\infty} f_p(\alpha_p)\,\mathrm{d}\alpha_p \end{aligned} \tag{64}$$

upon integration by parts. The first term vanishes and since the whole integral is unity

$$\overline{(\alpha_p - \alpha_p^0)\mathrm{d}\sigma_p/\mathrm{d}\alpha_p} = -k \tag{65}$$

Similarly, it can be shown

$$\overline{(\alpha_q - \alpha_q^0)\partial\sigma_{1\cdots n}/\partial\alpha_p} = \begin{cases} -k; & p = q \\ 0; & p \neq q \end{cases} \tag{66}$$

These averages involving the fluctuations $(\alpha_p - \alpha_p^0)$ are important in the subsequent derivations.

3. The regression of fluctuations

From an empirical point of view, the initial thermodynamic state has been employed to predetermine the course of an irreversible process according to definite laws, such as those for the flows of heat, electricity, and matter. Since molecular theory cannot completely define a thermodynamic state, it cannot completely predetermine an irreversible process. Nevertheless, the predictions of irreversible process from empirical laws may be interpreted statistically and with practical certainty from averages of a large number of cases of processes starting from the same initial state of thermodynamic equilibrium.

Now, consider the fluctuations of the variables of an isolated system over a long period of time. Whenever $\alpha_1, \alpha_2, \cdots \alpha_n$ have the values $\alpha_1', \alpha_2', \cdots \alpha_n'$, we record their values τ seconds later, and denote their averages by

$$\bar{\alpha}_1(\tau, \alpha_1', \cdots \alpha_n'), \cdots \bar{\alpha}_n(\tau, \alpha_1', \cdots \alpha_n') \tag{67}$$

Almost every time when $\alpha_1 = \alpha_1', \cdots \alpha_n = \alpha_n'$, the system will be in the phenomenological state

$$\Gamma'(1 \cdots n) = \Gamma(\alpha_1', \cdots \alpha_n') \tag{68}$$

and the average course of an irreversible process following this state, and described by the averages

$$\bar{\alpha}_1'(\tau, \Gamma_{1\cdots n}'), \cdots \bar{\alpha}_n(\tau, \Gamma_{1\cdots n}') \tag{69}$$

will be known from macroscopic experiments and may be regarded as properties of the state $\Gamma_{1\cdots n}'$. The "normal" properties corresponding to $\alpha_1', \cdots \alpha_n'$ of the fluctuating quantities $\alpha_1, \cdots \alpha_n$ are certainly those of the state $\Gamma_{1\cdots n}'$. Whether these "normal" and average values are interchangeable must be determined from consideration of each case. If this interchange is assumed, the relation

$$\bar{\alpha}_i(\tau, \alpha_1', \cdots \alpha_n') = \bar{\alpha}_i(\tau, \Gamma_{1\cdots n}'), \qquad i = 1, \cdots n \tag{70}$$

may be employed as a rule for predicting the average regression of fluctuations from the laws of irreversible processes.

4. *The derivation of reciprocal relations from the theory of microscopic reversibility applied to fluctuations*

We define the function A_{ji} by the equation

$$A_{ji}(\tau) \equiv \overline{\alpha_j(t)\alpha_i(t + \tau)} = \overline{\alpha_j'(t)\alpha_i(\tau, \alpha_j')} \tag{71}$$

Thus, whenever $\alpha_j = \alpha_j'$, we record α_i τ seconds later and denote it by $\alpha_i(t + \tau)$· $A_{ji}(\tau)$ may also be defined by the time integral

$$A_{ji}(\tau) \equiv \lim_{t'' \to \infty} \frac{1}{t'' - t'} \int_{t'}^{t''} \alpha_j(t)\alpha_i(t + \tau)\, dt \tag{72}$$

We shall now assume that the variables $\alpha_1, \alpha_2, \cdots \alpha_n$ represent deviations from the state of equilibrium in which their average values $\bar{\alpha}_1, \cdots \bar{\alpha}_n$ as well as their normal values α_j^0 vanish, or

$$\bar{\alpha}_i = \alpha_i^0 = 0, \qquad (i = 1, \cdots n) \tag{73}$$

This will be consistent with the evaluation of the averages given by equations 65 and 66, which depend upon the Boltzmann equation.

Now, if α_j and α_i are suitable "reversible" variables, the principle of microscopic reversibility requires that

$$A_{ji} = \overline{\alpha_j(t)\alpha_i(t + \tau)} = \overline{\alpha_i(t)\alpha_j(t + \tau)} = A_{ij} \tag{74}$$

This means that the event $\alpha_j = \alpha_j'$ followed τ seconds by $\alpha_i(t + \tau)$ occurs just as frequently as the event $\alpha_i = \alpha_j'$ followed τ seconds later by $\alpha_j(t + \tau)$.

Now, consider the type of irreversible process starting from the state $\Gamma_{1\cdots n}$ which can be described by the linear differential equations

$$\frac{d\bar{\alpha}_i}{dt} = \dot{\alpha}_i = \sum_{r=1}^{n} G_{ir} \frac{\partial \sigma_{1\cdots n}(\alpha_1, \alpha_2, \cdots \alpha_n)}{\partial \alpha_r}; \qquad (i = 1, \cdots n) \tag{75}$$

previously expressed by equation 51. Introducing relation 70 between average and normal values, we obtain

$$\bar{\alpha}_i(\tau, \alpha_j') = \bar{\alpha}_i(\tau, \Gamma_j') \tag{76}$$

in which Γ_j' is the state of maximum entropy for a given value of α_j', characterized by the relations

$$\alpha_j = \alpha_j' \tag{77}$$

$$\frac{\partial\sigma_{1\cdots n}}{\partial\alpha_r} = 0; \qquad (r \neq j) \tag{78}$$

and consequently,

$$\frac{\partial\sigma_{1\cdots n}}{\partial\alpha_j} = \left[\frac{d\sigma_j}{d\alpha_j}\right]_{\alpha_j=\alpha_j'} \tag{79}$$

For a short time integral, equation 75 leads to

$$\begin{aligned} \bar{\alpha}_i(\Delta t, \alpha_j') &= \bar{\alpha}_i(0, \alpha_j') + \dot{\alpha}_i \Delta t \\ &= \bar{\alpha}_i(0, \alpha_j') + \sum_{r=1}^{n} G_{ir} \frac{\partial\sigma_{1\cdots n}}{\partial\alpha_r} \Delta t \end{aligned} \tag{80}$$

whence from equations 77, 78, and 79:

$$\bar{\alpha}_i(\Delta t, \alpha_j') = \bar{\alpha}_i(0, \alpha_j') + G_{ij} \frac{d\sigma(\alpha_j')}{d\alpha_j} \Delta t \tag{81}$$

Therefore, $A_{ji}(\Delta t)$ (equation 71) may be written

$$A_{ji}(\Delta t) = \overline{\alpha_j(t)\alpha_i(t + \Delta t)} = \overline{\alpha_j'\alpha_i(0, \alpha_j')} + G_{ij}\Delta t \overline{\alpha_j \frac{d\sigma_j}{d\alpha_j}} \tag{82}$$

Since α_j is a small fluctuation from the equilibrium, it corresponds to $\overline{(\alpha_p - \alpha_p^0)}$ in equation 65 and consequently $\alpha_j d\sigma_j/d\alpha_j = -k$. Hence,

$$A_{ji}(\Delta t) = A_{ji}(0) - k\Delta t G_{ij} \tag{83}$$

and, similarly,

$$A_{ij}(\Delta t) = A_{ij}(0) - k\Delta t G_{ji} \tag{84}$$

The condition of microscopic reversibility, $A_{ji} = A_{ij}$, leads immediately to the general reciprocal relation

$$G_{ji} = G_{ij} \tag{85}$$

From this derivation, the importance of considering fluctuations from the thermodynamic equilibrium state becomes clear. The use of the Boltzmann relation leads to the simple introduction of the constant k in the above relations.

It should now be emphasized that these general reciprocal relations apply to systems whose macroscopic laws may be derived from initial thermodynamic conditions. Consequently, they are generally valid for laminar flow of liquids, conductance of heat, conductance of electricity, and the flow of matter in which no turbulence occurs. They are not valid for irreversible processes involving

Coriolis forces and the presence of external magnetic fields where it is known that the principle of dynamic reversibility no longer holds.

5. *The principle of the least dissipation of energy. The dissipation function*

It is now possible to define a dissipation function for irreversible processes which has a statistical significance similar to that of the entropy. Equation 75 may be written

$$\frac{\partial \sigma_{1\cdots n}(\alpha_1, \alpha_2, \cdots \alpha_n)}{\partial \alpha_i} = \sum_{r=1}^{n} \rho_{ir}\dot{\alpha}_r\,; \qquad (i = 1, 2, \cdots n) \tag{86}$$

where ρ_{ij} and G_{ij} are related by the well-known equation

$$\sum_{r=1}^{n} \rho_{ir} G_{rj} = \sum_{r=1}^{n} G_{ir}\rho_{rj} = \delta_{ij} = \begin{cases} 1, & (i = j) \\ 0, & (i \neq j) \end{cases} \tag{87}$$

The coefficients ρ_{ij} form the inverse matrix of G_{ij}. The symmetry conditions expressed by equation 85 become

$$\rho_{ij} = \rho_{ji}; \qquad (i = 1, 2, \cdots n) \tag{88}$$

We now define the *dissipation function* by

$$\Phi(\dot{\alpha}, \dot{\alpha}) \equiv \tfrac{1}{2}\sum_{i,j} \rho_{ij}\dot{\alpha}_i\dot{\alpha}_j \tag{89}$$

Utilizing the symmetry condition, $\rho_{ij} = \rho_{ji}$, we find from equations 86 and 89 that

$$\frac{\partial \sigma_{1\cdots n}(\alpha_1, \cdots \alpha_n)}{\partial \alpha_i} = \frac{\partial \Phi(\dot{\alpha}, \dot{\alpha})}{\partial \dot{\alpha}_i} \tag{90}$$

Further, we define $\dot{S}(\alpha, \dot{\alpha})$ by

$$\dot{S}(\alpha, \dot{\alpha}) \equiv \sum_{r=1}^{n} \frac{\partial \sigma_{1\cdots n}}{\partial \alpha_r}\dot{\alpha}_r \tag{91}$$

for the rate of increase of entropy.

If we adopt the convention that the velocities, $\dot{\alpha}_1, \dot{\alpha}_2, \cdots \dot{\alpha}_n$, are the only variables, we may obtain a variation principle in the form

$$\delta[\dot{S}(\alpha, \dot{\alpha}) - \Phi(\dot{\alpha}, \dot{\alpha})] = 0 \tag{92}$$

for according to equation 90,

$$\delta[\dot{S}(\alpha, \dot{\alpha}) - \Phi(\dot{\alpha}, \dot{\alpha})] = \sum_{i=1}^{n} \left(\frac{\partial \sigma_{1\cdots n}}{\partial \alpha_i} - \frac{\partial \Phi}{\partial \dot{\alpha}_i}\right)\delta\dot{\alpha}_i = 0 \tag{93}$$

This is a generalization of Lord Rayleigh's "principle of the least dissipation of energy." The dissipation function according to equations 86, 87, and 88 equals one-half the rate of increase of entropy

$$2\Phi(\dot{\alpha}, \dot{\alpha}) = \dot{S}(\alpha, \dot{\alpha}) \tag{94}$$

and

$$\Phi(\dot{\alpha}, \dot{\alpha}) = \frac{1}{2} \sum_{i,j} \rho_{ij} \dot{\alpha}_i \dot{\alpha}_j = \frac{1}{2} \sum_i \dot{\alpha}_i \frac{\partial \Phi}{\partial \dot{\alpha}_i} \tag{95}$$

Since according to thermodynamics $\dot{S} \geqq 0$, $\Phi(\dot{\alpha}, \dot{\alpha})$ is positive, and therefore the variation given by equation 92 is the condition for a maximum. Thus,

$$\dot{S}(\alpha, \dot{\alpha}) - \Phi(\dot{\alpha}, \dot{\alpha}) = \text{maximum} \tag{96}$$

This quantity is similar to the entropy in reversible processes, and it is clear from the preceding discussion that it depends on the existence of reciprocal relations.

6. *Condensed summary and formulation*

We have noted that the quantities $\partial\sigma/d\alpha$ as expressed by equations 44, 45, and 46 correspond to "the forces" $\mathbf{X}_1$, $\mathbf{X}_2$, and $\mathbf{X}_3$ in the phenomenological equations with which we introduced the subject, and that $\dot{\alpha}_r$ corresponds to the "flows" $\mathbf{J}_1$, $\mathbf{J}_2$, $\mathbf{J}_3$. If ρ_{ir} be replaced by R_{ir}, then equation 86 for three processes may be written in the expanded form

$$\begin{aligned}\mathbf{X}_1 &= R_{11}\mathbf{J}_1 + R_{12}\mathbf{J}_2 + R_{13}\mathbf{J}_3\\ \mathbf{X}_2 &= R_{21}\mathbf{J}_1 + R_{22}\mathbf{J}_2 + R_{23}\mathbf{J}_3\\ \mathbf{X}_3 &= R_{31}\mathbf{J}_1 + R_{32}\mathbf{J}_2 + R_{33}\mathbf{J}_3\end{aligned} \tag{97}$$

where R_{11} etc. are related to L_{11} etc. in equation 23 by the transformation equations

$$\sum_{m=1}^{3} L_{im}R_{mk} = \sum_{m=1}^{3} R_{im}L_{mk} = \delta_{ik} = \begin{cases}1, & i = k\\ 0, & i \neq k\end{cases} \tag{98}$$

and where the reciprocal relations

$$R_{12} = R_{21}; \qquad R_{13} = R_{31}; \qquad R_{32} = R_{23} \tag{99}$$

$$L_{12} = L_{21}; \qquad L_{13} = L_{31}; \qquad L_{32} = L_{23} \tag{100}$$

are valid. The dissipation function becomes

$$2\Phi(\mathbf{J}, \mathbf{J}) = \frac{1}{T} \sum_{i,k} R_{ik}\mathbf{J}_i\mathbf{J}_k \tag{101}$$

Although these symmetry conditions have not been tested experimentally, the preceding argument and experiments on thermoelectric phenomena, cells with liquid junction, etc. lend confidence to their fundamental correctness. If these reciprocal relations are valid for a given combination of processes, then the dissipation function may be expressed by equation 101 and the function $\dot{S}(\alpha, \dot{\alpha}) - \Phi(\dot{\alpha}, \dot{\alpha})$ will be a maximum as required by equation 96. For the process of diffusion in a system of any number of components, the reciprocal relation $\mathfrak{M}_{ik} = \mathfrak{M}_{ki}$ (equation 26) may be assumed with confidence. In Section IV we shall derive a detailed formula for the dissipation function for combined conductance and diffusion of a single electrolyte in dilute solution. Onsager and Fuoss (80) have treated the more general case of systems containing more than two components. These relations are indeed important, although no direct determination of the dissipation function has been made up to the present time.

A general statistical theory of transport processes which goes further than the treatment of Onsager has been developed recently by Kirkwood (51a). A generalized treatment of Brownian motion is obtained in which relations between friction constants and intermolecular forces are deduced for the first time. Important progress in this direction is to be anticipated. Another contribution by Leaf (57a), who has succeeded in obtaining a general phenomenological theory of transport processes in liquids from the equations of thermodynamics and hydrodynamics, deserves serious consideration. Expressions for the flow of entropy and the flows of masses as well as reciprocal relations between the coefficients in these processes are derived.

IV. THEORY OF ELECTROLYTIC DIFFUSION

1. General considerations and force relations (80)

The quantitative theory of electrolytes depends on two fundamental facts. (*1*) To a first approximation, the properties of electrolytic solutions are additive and to a great extent they are functions of the individual ions. On this fact depend the law of Kohlrausch, of the additivity of ionic conductances at infinite dilution, and the law of Nernst (71), which permits the computation of the diffusion coefficients of electrolytes at infinite dilution from electrical properties derived from other measurements. These behaviors and the fact that deviations from additivity in electrolytic solutions increase slowly with increasing concentration of solute are advantages which electrolytic solutions possess and which are not characteristic of molecular solutions.

(*2*) The deviations from additivity may be explained by the interactions between ions as charged particles and between ions and the solvent molecules. Fortunately, the law of attraction between charged particles is known, which makes possible a highly developed theory of the properties of ionic solutions of sufficient dilution so that solvent interaction is a higher order effect. In concentrated solutions, little progress has been achieved towards a detailed theory nor may we expect quantitative results until we know the structure of the solvent and the laws of the forces between all the constituents of the solution.

A treatment of the motion of ions in homogeneous media and fields of forces sufficiently general to include a basis for a detailed theory of conductance and diffusion of electrolytes may be expressed in a simple form. For, suppose the forces $\mathbf{k}_1, \mathbf{k}_2, \cdots \mathbf{k}_s$ act on the ions, 1, 2, $\cdots s$, and that these are balanced by necessity by a compensating force $\mathbf{k}_0$ acting on the molecules of the solvent. Then

$$n_0\mathbf{k}_0 + \sum_{i=1}^{s} n_i\mathbf{k}_i = 0 \qquad (102)$$

This system of forces is produced by a combination of electrical fields, $-\nabla e_i\psi$ acting on each ion, where ψ is the electrical potential, with the concentration gradients, ∇n_1, ∇n_2, $\cdots$ ∇n_s. The required condition of electrical neutrality is

$$\sum_{i=1}^{s} e_i \nabla n_i \equiv \nabla \sum_{i=1}^{s} n_i e_i = \nabla(0) = 0 \tag{103}$$

That these relations are consistent with thermodynamics is proved by the following reasoning. Equilibrium may be maintained with these electrical forces and gradients by superimposed forces of opposite signs and of different kinds, $-\mathbf{k}_0$, $-\mathbf{k}_1$, $\cdots$ $-\mathbf{k}_s$ designated by

$$\mathbf{k}_i = -\nabla\bar{\mu}_i = -\nabla(\mu_i + e_i\psi) \tag{104}$$

where μ_i is the Gibbs chemical potential. Here we employ the quantity $\bar{\mu}_i$, which Brønsted (7) and Guggenheim (38) denoted the "electrochemical potential." In addition to these relations, thermodynamics imposes the Gibbs-Duhem equation:

$$\sum_{i=0}^{s} n_i \, d\mu_i = 0 \tag{105}$$

From equations 102, 103, and 104 we obtain

$$n_0 \nabla\mu_0 + \sum_{i=1}^{s} n_i \nabla\mu_i + \nabla \sum_{i=1}^{s} n_i e_i \psi = 0 \tag{106}$$

whence by equation 105

$$\nabla \sum_{i=1}^{s} n_i e_i \psi = 0 \tag{107}$$

and consequently

$$\nabla \sum_{i=1}^{s} n_i e_i = 0 \tag{108}$$

the required condition of electrical neutrality.

2. *The limiting value of the diffusion coefficient of a simple electrolyte. Extension of Nernst's equation*

We shall now consider the case of diffusion of a simple electrolyte dissociating into ν_1 cations and ν_2 anions and omit all the effects of interactions between ions. In diffusion, the condition of the absence of electric current requires that the ions migrate with the same velocity:

$$\mathbf{v} = \mathbf{v}_1 = \mathbf{v}_2 \tag{109}$$

The flows are given by

$$\mathbf{J}_1 = n_1\mathbf{v}_1; \qquad \mathbf{J}_2 = n_2\mathbf{v}_2 \tag{110}$$

Further, the velocities are related to the forces and the electrochemical potentials according to

$$\mathbf{v} = \mathbf{k}_1\omega_1 = \mathbf{k}_2\omega_2 = -\omega_1\nabla\bar{\mu}_1 = -\omega_2\nabla\bar{\mu}_2 \tag{111}$$

where ω_1 and ω_2 are the mobilities, or velocities in unit force fields. Now

$$\bar{\mu}_1 + \bar{\mu}_2 = \mu_1 + \mu_2 + \psi(e_1 + e_2) = \mu_1 + \mu_2 \tag{112}$$

since $e_1 = -e_2$ and their sum is zero. Therefore we write

$$\mathbf{k} = \nu_1\mathbf{k}_1 + \nu_2\mathbf{k}_2 = -\nabla\mu \tag{113}$$

From these equations

$$\mathbf{k}_2 = \mathbf{k}_1\frac{\omega_1}{\omega_2} = \mathbf{k}_1\frac{\rho_2}{\rho_1} \tag{114}$$

where $\rho = 1/\omega$ denotes a coefficient of friction. From the last two relations

$$\mathbf{k}_1 = -\frac{\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}\nabla\mu = -\frac{\rho_1}{\nu_1\rho_1 + \nu_2\rho_2}\nabla\mu \tag{115}$$

and

$$\mathbf{v} = -\frac{\omega_1\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}\nabla\mu \tag{116}$$

If the solution is ideal

$$\mu = \mu^0 + (\nu_1 + \nu_2)\,kT\ln n$$

where μ^0 represents the standard state of chemical potential and n is the number of mols of electrolyte per cubic centimeter. Then

$$\nabla\mu = (\nu_1 + \nu_2)kT\frac{\nabla n}{n} \tag{117}$$

and

$$\mathbf{J} = n\mathbf{v} = -\frac{(\nu_1 + \nu_2)\omega_1\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}kT\nabla n = -\mathfrak{D}\nabla n \tag{118}$$

Consequently, the diffusion coefficient is given by

$$\mathfrak{D}_0 = \mathfrak{D}_{n\to 0} = \frac{(\nu_1 + \nu_2)\omega_1\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}kT \tag{119}$$

which is the famous limiting equation of Nernst (71) valid at infinite dilution of electrolyte.

If we introduce the real value of μ, then

$$\nabla\mu = (\nu_1 + \nu_2)kT\,\nabla\ln\zeta_{\pm}n \tag{120}$$

$$\mathbf{J} = n\mathbf{v} = -\frac{\omega_1\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}(\nu_1 + \nu_2)kT\left(1 + n\frac{\partial\ln\zeta_{\pm}}{\partial n}\right)\nabla n \tag{121}$$

and

$$\mathfrak{D} = \frac{(\nu_1 + \nu_2)\omega_1\omega_2}{\nu_1\omega_2 + \nu_2\omega_1}kT\left(1 + n\frac{\partial\ln\zeta_{\pm}}{\partial n}\right) \tag{122}$$

or according to equation 6

$$\mathfrak{D} = (\nu_1 + \nu_2)\frac{\mathfrak{M}}{n}kT\left(1 + n\frac{\partial \ln \zeta_\pm}{\partial n}\right) \tag{123}$$

Thus, if the variation of $\mathfrak{D}$ with concentration depended entirely on the term in parenthesis representing the deviation from the ideal solution, $\mathfrak{M}/n$ would be constant. This is not the case, nor is it to be expected on theoretical grounds because the velocities of the ions in diffusion will be affected by various interactions. $\mathfrak{M}/n$ is seen to be a mobility term, or a velocity under unit potential gradient caused by a force $\mathbf{k} = -\nabla\mu$. We shall now undertake the theoretical evaluation of $\mathfrak{M}/n$.

3. Elements of the interionic attraction theory of Debye required for the theory of diffusion

In an ionic solution, the electrostatic forces tend to bring the ions of opposite sign together while the thermal force acts contrariwise. Combination of these two effects was shown by Debye (17, 18) to lead to the formation of ionic atmospheres whereby each positive ion, by inducing a negative charge density in its neighborhood, will be surrounded by an "atmosphere" which contains on the average more negative and less positive ions than the bulk of the solution. Similarly, a negative ion will be surrounded by a positive atmosphere.

When the system is undisturbed or at equilibrium, the potential $\psi_j^0(r)$ of the central ion, j, and its atmosphere at a distance r from it is given by

$$\psi_j^0(r) = Ae^{-\kappa r}/r \tag{124}$$

where A is an integration constant and κ is defined by

$$\kappa^2 \equiv \frac{4\pi\epsilon^2}{DkT}\sum_{i=1}^{s} n_i z_i^2 \tag{125}$$

This quantity has the dimensions of a reciprocal distance, and $1/\kappa$ is related to the potential of the ionic atmosphere ψ_j^{0*} in exactly the same way as the distance r from a charge e is related to the potential produced at this distance. Here ϵ is the electronic charge, D the dielectric constant, z_i the valence of the ions surrounding the j ion, and k the gas constant per molecule. The summation is over all the kinds of ions in the solution. It is to be noted that z_i carries the sign of the charge. We shall use the customary symbolism of vertical bars $|z_i|$ when magnitude is represented.

The constant A in equation 124 is found to be

$$A = \frac{e_j}{D}\frac{e^{\kappa a}}{1 + \kappa a} \tag{126}$$

where a is the parameter which represents the mean distance of approach of ions positive or negative. By substituting this value of A in equation 124, we obtain

$$\psi_j^0(r) = \frac{e_j}{D} \frac{e^{\kappa a} e^{-\kappa r}}{(1 + \kappa a) r} \tag{127}$$

This is one of the relations required for our subsequent derivations (41).

We shall also use the Maxwell-Boltzmann equation for the number of ions of one kind in the neighborhood of another ion, namely:

$$n_{ji} = n_i e^{-(e_i \psi_j^0 / kT)} \simeq n_i \left[1 - \frac{e_i \psi_j^0}{kT} + \frac{1}{2} \left(\frac{e_i \psi_j^0}{kT} \right)^2 - \cdots \right] \tag{128}$$

We observe that $\psi_j^0 e_i$ is the potential energy, the work required to bring an i ion of charge e_i from infinity to a distance r from the j ion in a field derivable from the potential ψ_j^0.

For the computation of the activity coefficient term in an equation corresponding to equation 123 we shall use the relation

$$\log y_\pm = -\frac{\mathfrak{S}_{(f)} \sqrt{c}}{1 + A' \sqrt{c}} + B2c + D'c^2 - \log \left[\frac{d + 0.001c[2M_1 - M_2]}{d_0} \right] \tag{129}$$

where $y_\pm$ is the mean activity coefficient of electrolyte on the mol per liter scale, $\mathfrak{S}_{(f)}$ is the Debye and Hückel constant, A' is the parameter containing the mean distance of approach of the ions, B and D' are empirical constants, and c is in mols per liter (41). The last term on the right is the relation required for converting $\log y_\pm$ to $\log f_\pm$, where $f_\pm$ is the "rational" activity coefficient. In this term d is the density of the solution, d_0 the density of the solvent, M_1 the molecular weight of electrolyte, and M_2 that of solvent (41).

4. Preliminary theory of the action of forces which cause the migration of ions

When the ions of a system in thermodynamic equilibrium are subjected to certain forces, the ion atmospheres are disturbed. If the perturbing force is removed, the displaced ions will return to their original state of distribution. This will require a finite time which is denoted the time of relaxation τ and is expressed by

$$\tau = \frac{1}{\kappa^2 kT \overline{1/\rho}} \tag{130}$$

where

$$\overline{1/\rho} = \frac{\sum_{i=1}^{s} n_i e_i^2 / \rho_i}{\sum_{i=1}^{s} n_i e_i^2} \tag{131}$$

If we let $\rho_1 = \rho_2 = \cdots \rho_i$, then

$$\tau = \frac{\rho_i}{\kappa^2 kT} \tag{132}$$

simplified result which is sufficient for our purpose. Comparison with the erivations (Section IV, 2) will show that kT/ρ_i is really the diffusion coefficient f an ion.

The action of forces which cause the ions to move leads to two effects which are subject to quantitative treatment. The first of these is the "time of relaxation" effect, whereby the applied force causes the ion to move away from its atmosphere. This ion will attract the lagging atmosphere and cause the original spherical distribution of the ions in the atmosphere to assume an assymetry. The net effect will be the retardation of the motion of the ion. The time of relaxation effect has been treated in detail by Onsager (80). The second effect, that of electrophoresis, is due to the fact that the ion migrates in one direction while its atmosphere and the solvent move in the opposite direction. This counterwise motion has the effect of retarding the motion of the ion.

5. The time-of-relaxation effect

Under the influence of a force $\mathbf{k}_j$ an ion will migrate with a velocity

$$\mathbf{v}_j = \frac{\mathbf{k}_j}{\rho_j} \tag{133}$$

When a force is applied, the ion will move away from its atmosphere and will be ahead of its lagging atmosphere by a distance

$$\mathbf{v}_j\tau = \frac{\tau\mathbf{k}_j}{\rho_j} \tag{134}$$

The ratio of this distance to $1/\kappa$ (the radius of the atmosphere), or $\tau\mathbf{k}_j\kappa/\rho_j$, will be a measure of the asymmetry of the atmosphere. The directed force $\Delta\mathbf{k}_j$ due to this effect will be given roughly by multiplying this quantity by the Coulombic force between the ion and its atmosphere, $e_j^2\kappa^2/D$, whence

$$\Delta\mathbf{k}_j = \frac{-e_j^2\kappa^2}{D}\,\frac{\kappa\tau}{\rho_j}\,\mathbf{k}_j = -\frac{e_j^2\kappa\mathbf{k}_j}{DkT} \tag{135}$$

after eliminating τ by means of equation 132. This result differs from Onsager's final result by a numerical factor which can only be computed by an elaborate method. Its derivation, however, shows the essential characteristics of the theory.

It is sufficient for our purpose of calculating the dissipation function for simultaneous diffusion and electrical conductance of ions to omit the complicated calculation of the effect of a superimposed electrical field and simply state the required results of the theory. If $\mathbf{X}$ is the x-component of the external electrical field, then $\Delta\mathbf{X}$, the field caused by the asymmetry of the ionic atmospheres for an electrolyte dissociating into two ions, is given by

$$\Delta\mathbf{X} = \Delta\mathbf{X}_1 = \Delta\mathbf{X}_2 = \frac{e_1e_2\kappa}{3DkT}\left(\frac{q^*}{1+\sqrt{q^*}}\right)\mathbf{X} \tag{136}$$

where

$$q^* = \frac{n_1e_1^2\omega_1 + n_2e_2^2\omega_2}{(n_1e_1^2 + n_2e_2^2)(\omega_1 + \omega_2)} \tag{137}$$

The forces $\mathbf{k}_1$ and $\mathbf{k}_2$ are equal to $\mathbf{X}e_1$ and $\mathbf{X}e_2$, respectively, so that $\mathbf{X}$ may be expressed by

$$\mathbf{X} = \frac{\omega_2 \mathbf{k}_2 - \omega_1 \mathbf{k}_1}{e_2 \omega_2 - e_1 \omega_1} \tag{138}$$

Upon elimination of $\mathbf{X}$ and rearranging we readily obtain

$$\Delta \mathbf{X}_1 = \Delta \mathbf{X}_2 = \frac{e_1 e_2 \kappa}{3DkT} \frac{(1 - \sqrt{q^*}) n_2 e_2}{n_1 e_1^2 \omega_2 + n_2 e_2^2 \omega_1} (\omega_2 \mathbf{k}_2 - \omega_1 \mathbf{k}_1) \tag{139}$$

the relation we shall use later to compute the dissipation function.

From this equation a very important conclusion in relation to the theory of diffusion is apparent. In the process of diffusion of a simple electrolyte, both kinds of ions travel with the same velocity and in the same direction. Thus

$$\mathbf{v}_1 = \omega_1 \mathbf{k}_1 = \omega_2 \mathbf{k}_2 = \mathbf{v}_2 \tag{140}$$

whence by equation 139 the *time-of-relaxation effect vanishes and the ionic atmospheres retain their original spherical symmetry.* In conductance, not only do the ionic velocities differ but positive and negative ions move in opposite directions, which leads to an additive contribution to the current.

6. *Electrophoresis*

For conductance, where the force $\mathbf{k}_j$ is proportional to the charge on an ion, e_j,

$$\mathbf{k}_0 = 0; \quad \mathbf{k}_j = e_j \mathbf{X}; \quad j = 1, 2, \cdots s \tag{141}$$

the effect of electrophoresis is easily computed. The ionic atmosphere will possess a charge, $-e_j$, and will be subjected to the force $-e_j\mathbf{X}$, which will cause the atmosphere and the solvent containing it to migrate in a direction opposite to the motion of the ion. Therefore the ion moves in a counter-current of solvent which has the net effect of retarding its velocity. As pointed out by Debye (18), the velocity of the counter-current may be computed by assuming that the entire charge, $-e_j$, is at a distance $1/\kappa$ from the central ion and distributed over a sphere of radius $1/\kappa$. Further, it is assumed that Stokes law applies to the motion of this sphere and therefore the velocity of the electrophoretic effect is given by

$$\Delta \mathbf{v}_j = -\frac{\mathbf{k}_j \kappa}{6\pi\eta} = -\frac{\mathbf{X} e_j \kappa}{6\pi\eta} \tag{142}$$

where η is the viscosity of the medium.

The velocity of the ion during electrical current flow under a field $\mathbf{X}$ may be expressed in the following manner. If the ion possessed no atmosphere, its velocity in a field $\mathbf{X}$ would be $\mathbf{X}e_j\omega_j$; but owing to its lagging atmosphere it is subjected to an additional force in the opposite direction, $-\omega_j \Delta \mathbf{k}_j$, and the electrophoretic effect $\Delta \mathbf{v}_j$. Therefore, the net velocity becomes

$$\mathbf{v}_j = \omega_j \mathbf{X} e_j - \omega_j \Delta \mathbf{k}_j - \frac{\mathbf{X} e_j \kappa}{6\pi\eta} \tag{143}$$

When an electrolyte diffuses into a solvent, both ions travel with the same velocity, a phenomenon required by electrical neutrality. In this case, electrophoresis occurs because as the ions move in one direction, they take positions previously occupied by the water molecules which in a system at constant volume must move in the opposite direction. The calculation of the electrophoretic effect will require therefore an estimate of the volume force acting on the water molecules.

According to equation 102

$$n_0\mathbf{k}_0 + \sum_{i=1}^{s} n_i\mathbf{k}_i = 0 \tag{144}$$

the forces acting on the ions must be balanced by a compensating force acting on the water molecules. For simplicity, we shall let the subscript σ represent summation. Then

$$n_0\mathbf{k}_0 + n_\sigma\mathbf{k}_\sigma = 0 \tag{145}$$

Now, the presence of a j ion influences the average concentration of the ions in a given volume element dV at a distance r from it. The volume force becomes $n_{j\sigma}\mathbf{k}_\sigma\,\mathrm{d}V$ and the net force on the volume element will be

$$(n_{j\sigma}\mathbf{k}_\sigma + n_0\mathbf{k}_0)\mathrm{d}V = (n_{j\sigma} - n_\sigma)\mathbf{k}_\sigma\mathrm{d}V \tag{146}$$

In spherical coördinates, the force acting on a spherical shell at the distance, r, from the j ion will be

$$\mathcal{F}\mathrm{d}r = 4\pi r^2(n_{j\sigma} - n_\sigma)\mathbf{k}_\sigma\mathrm{d}r \tag{147}$$

and is distributed over the shell of radius r. This force will cause the points within the sphere to move with a velocity, $\mathbf{v}$, given by Stokes law:

$$\mathbf{v} = \frac{\mathcal{F}}{6\pi\eta r} \tag{148}$$

Now $(n_{ji} - n_i)$ may be obtained from the Boltzmann equation (equation 128) and ψ_j^0 may be eliminated from this expression by means of equation 127. Substituting this result in equation 147, the force acting on the spherical shell is found to be

$$\mathcal{F}\,\mathrm{d}r = 4\pi r^2\left\{-\frac{e_je_\sigma}{DkT}\,\frac{e^{\kappa a}}{(1+\kappa a)}\,\frac{e^{-\kappa r}}{r} + \frac{1}{2}\left(\frac{e_je_\sigma}{DkT}\,\frac{e^{\kappa a}}{(1+\kappa a)}\right)^2\frac{e^{-2\kappa r}}{r^2}\right\}n_\sigma\mathbf{k}_\sigma\,\mathrm{d}r \tag{149}$$

or more simply

$$\mathcal{F}\mathrm{d}r = 4\pi(-A_1re^{-\kappa r} + A_2e^{-2\kappa r})\,\mathrm{d}r \tag{150}$$

Upon substituting in Stokes equation

$$\mathrm{d}\mathbf{v}_j = \frac{\mathcal{F}\,\mathrm{d}r}{6\pi\eta r} = \frac{2}{3\eta}\left(-A_1e^{-\kappa r} + \frac{A_2e^{-2\kappa r}}{r}\right)\mathrm{d}r \tag{151}$$

is obtained and upon integration from a to ∞, the total velocity $\Delta\mathbf{v}_j$ is found to be

$$\Delta\mathbf{v}_j = \frac{2}{3\eta}\left(-\frac{A_1 e^{-\kappa a}}{\kappa} + A_2 Ei(2\kappa a)\right) \tag{152}$$

where $Ei(2\kappa a)$ is the well-known exponential integral function

$$Ei(x) = \int_x^\infty e^{-t}\frac{dt}{t} = -0.5772 - \ln x + x - \frac{x^2}{2\underline{|2}} + \cdots \tag{153}$$

Introducing A_1 and A_2, the final result is

$$\Delta\mathbf{v}_j = -\frac{2}{3\eta}\frac{e_j e_\sigma n_\sigma \mathbf{k}_\sigma}{DkT\kappa(1+\kappa a)} + \frac{1}{3\eta}\left(\frac{e_j}{DkT}\frac{e_\sigma e^{\kappa a}}{(1+\kappa a)}\right)^2 n_\sigma \mathbf{k}_\sigma Ei(2\kappa a) \tag{154}$$

At low concentrations when κa is negligible, and when the force, $\mathbf{k}_i = \mathbf{X}e_i$, is proportional to the charge,

$$\Delta\mathbf{v}_j = -\frac{2}{3\eta}\frac{e_j\mathbf{X}\sum_{i=1}^{s} n_i e_i^2}{DkT\kappa} \tag{155}$$

From the equation which defines κ, we find that

$$\sum_{i=1}^{s} n_i e_i^2 = \frac{DkT\kappa^2}{4\pi} \tag{156}$$

whence

$$\Delta\mathbf{v}_j = -\frac{1}{6}\frac{\mathbf{X}e_j\kappa}{\pi\eta} \tag{157}$$

which is the simple result for the electrophoretic effect previously obtained for the case of conductance.

It is well to note that in the derivation of equation 154, the Debye and Hückel calculation of the Coulombic force effect and their second approximation to include the effect of the mean distance of approach of the ions were employed. Consequently, this estimate of the electrophoretic effect can only be expected to approach validity at low concentrations of electrolytes. Deviations due to on–solvent and other short-range ionic interactions may be expected at higher concentrations and at the present time there exists no means of computing their nagnitudes.

7. *Extension of the theory of the diffusion of a simple electrolyte to include electrophoresis*

Now the effect of electrophoresis is to alter the forces, $\mathbf{k}_1$ and $\mathbf{k}_2$, acting upon he ions of an electrolyte by decreasing the velocities. Thus

$$\mathbf{k}_j = \rho_j(\mathbf{v} - \Delta\mathbf{v}_j); \quad j = 1, 2 \tag{158}$$

hich when introduced in equation 113 gives

$$-\nabla\mu = \mathbf{k} = \nu_1\rho_1(\mathbf{v} - \Delta\mathbf{v}_1) + \nu_2\rho_2(\mathbf{v} - \Delta\mathbf{v}_2) \tag{159}$$

For the electrophoretic terms, equation 154 yields

$$\Delta\mathbf{v}_j = \mathbf{v}\left[-\frac{2}{3\eta}(n_1e_1\rho_1 + n_2e_2\rho_2)\frac{e_j}{DkT\kappa(1+\kappa a)} + \frac{1}{3\eta}(n_1e_1^2\rho_1 + n_2e_2^2\rho_2)\left(\frac{e_j}{DkT}\right)^2\phi(\kappa a)\right] \tag{160}$$

where

$$\phi(\kappa a) = e^{2\kappa a}Ei(2\kappa a)/(1+\kappa a)^2 \tag{161}$$

Upon substituting the values of $\Delta\mathbf{v}_1$ and $\Delta\mathbf{v}_2$ in equation 159 and using the relations $n_1e_1 = -n_2e_2$ and $\nu_1e_1 = \nu_2e_2$, we obtain for the gradient of the chemical potential:

$$-\nabla\mu = \mathbf{v}\left[\nu_1\rho_1 + \nu_2\rho_2 + (\rho_1-\rho_2)^2\frac{\nu_1\nu_2}{\nu_1+\nu_2}\frac{\kappa}{6\pi\eta(1+\kappa a)} - \left(\frac{\nu_2\rho_1+\nu_1\rho_2}{\nu_1+\nu_2}\right)^2\frac{\kappa^4\phi(\kappa a)}{48\pi^2\eta n}\right] \tag{162}$$

$\nabla\mu$ is the force per molecule of solute when the concentration is expressed in molecules per cubic centimeter. The electrophoretic effect is seen to depend on the changes in frictional forces caused by the action of the Coulomb forces. The third term on the right is the first-order electrolytic effect which depends on the factor $\rho_1^2 - 2\rho_1\rho_2 + \rho_2^2$, whose terms correspond to cation–cation, cation–anion, and anion–anion effects, respectively. The last term of this equation is of the order $n \log n$.

If this equation be solved for $\mathbf{v}$ and multiplied by n, and higher order terms be discarded as negligible, we obtain the flow, $\mathbf{J}$. Thus

$$\mathbf{J} = n\mathbf{v} = -\mathfrak{M}\nabla\mu = -\left(\frac{n\omega_1\omega_2}{\nu_1\omega_2+\nu_2\omega_1} + \Delta\mathfrak{M}\right)\nabla\mu \tag{163}$$

where

$$\Delta\mathfrak{M} = -\left(\frac{\omega_1-\omega_2}{\nu_1\omega_2+\nu_2\omega_1}\right)^2\left(\frac{\nu_1\nu_2}{\nu_1+\nu_2}\right)\frac{\kappa n}{6\pi\eta(1+\kappa a)} + \left(\frac{\nu_1\omega_1+\nu_2\omega_2}{\nu_1\omega_2+\nu_2\omega_1}\right)^2\frac{1}{(\nu_1+\nu_2)^2}\frac{\kappa^4\phi(\kappa a)}{48\pi^2\eta} \tag{164}$$

It is now necessary to convert this equation into a form convenient for practical computations. To this end, we recall that

$$n = N\bar{n} = Nc/1000 \tag{165}$$

where $\bar{n}$ and c are the concentrations in mols per cubic centimeter and mols per liter, respectively. Further

$$\mu N = \bar{\mu} \tag{166}$$

and consequently

$$\mathbf{J} = \bar{n}\mathbf{v} = -\bar{\mathfrak{M}}\nabla\bar{\mu} = -\left(\frac{\mathfrak{M}}{N^2}\right)\nabla\mu \tag{167}$$

Now, the equivalent conductance of an ion, λ_i, is equal to the current at a potential gradient of one volt per centimeter produced by 1 gram-equivalent. Thus

$$\lambda_i = 96{,}500\ \bar{u}_i = 96{,}500\ |e_i|\ \omega_i \qquad (168)$$

If $\bar{u}_i$ is expressed in volts per centimeter and u_i in electrostatic units, then

$$\omega_i = u_i/|z_i|\epsilon = 300\ \bar{u}_i/|z_i|\epsilon \qquad (169)$$

$$= \frac{300\ \lambda_i}{96{,}500\ |z_i|\ \epsilon} \qquad (170)$$

Using this relation to eliminate ω_1 and ω_2 from equation 164, and taking into consideration the relations given by equations 165 and 166, we find that

$$\overline{\mathfrak{M}} = 1.0748 \times 10^{-20} \frac{\lambda_1^0 \lambda_2^0 c}{\nu_1 |z_1| \Lambda^0} + \Delta\overline{\mathfrak{M}}' + \overline{\mathfrak{M}}'' \qquad (171)^5$$

where the superscript signifies limiting equivalent conductance and $\Lambda^0 = \lambda_1^0 + \lambda_2^0$. The number of cations and anions into which an electrolyte dissociates are ν_1 and ν_2, respectively, and $|z_1|$ and $|z_2|$ are the magnitudes of their valences. Making use of the relations $\nu_1|z_1| = \nu_2|z_2|$, $\nu_2 = |z_1|$, and $\nu_1 = |z_2|$, $\Delta\overline{\mathfrak{M}}'$ and $\Delta\overline{\mathfrak{M}}''$ are given by

$$\Delta\overline{\mathfrak{M}}' = -\frac{(|z_2|\lambda_1^0 - |z_1|\lambda_2^0)^2}{|z_1 z_2|(\nu_1 + \nu_2)\Lambda^{02}} \frac{3.1322 \times 10^{-19}}{\eta_0 (DT)^{\frac{1}{2}}} \frac{c\sqrt{\Gamma}}{(1 + \kappa a)} \qquad (172)$$

$$\Delta\overline{\mathfrak{M}}'' = \left(\frac{z_2^2\lambda_1^0 + z_1^2\lambda_2^0}{\Lambda^0}\right)^2 \frac{9.304 \times 10^{-13}}{\eta_0 (DT)^2} c^2 \phi(\kappa a) \qquad (173)$$

The coefficient of diffusion in terms of these quantities is expressed by

$$\mathfrak{D} = \mathfrak{M} \frac{\partial \mu}{\partial n} = \overline{\mathfrak{M}} \frac{\partial \bar{\mu}}{\partial \bar{n}} = 1000\overline{\mathfrak{M}} \frac{\partial \bar{\mu}}{\partial c} \qquad (174)$$

and, since

$$c\frac{\partial \bar{\mu}}{\partial c} = (\nu_1 + \nu_2)RT\left(1 + c\frac{\partial \ln y_\pm}{\partial c}\right) \qquad (175)$$

we find that

$$\mathfrak{D} = (\nu_1 + \nu_2)1000RT\left(\frac{\overline{\mathfrak{M}}}{c}\right)\left(1 + c\frac{\partial \ln y_\pm}{\partial c}\right) \qquad (176)$$

Here, the activity coefficient is mean activity divided by the molar concentration.

For an electrolyte dissociating into two ions

$$\mathfrak{D} = 16.629 \times 10^{10}\, T \frac{\overline{\mathfrak{M}}}{c}\left(1 + c\frac{\partial \ln y_\pm}{\partial c}\right) \qquad (177)$$

---

[5] The numerical values in this and the following equations have been computed from the fundamental constants given by R. T. Birge (Rev. Modern Phys. 13, 233 (1941).

and equations 171, 172, and 173 reduce to a convenient form

$$\left(\frac{\mathfrak{M}}{c}\right)10^{20} = \frac{1.0748}{z}\left(\frac{\lambda_1^0\lambda_2^0}{\Lambda^0}\right) - \frac{0.4404}{\mathring{a}\eta_0}\left(\frac{\lambda_1^0 - \lambda_2^0}{\Lambda^0}\right)^2 \frac{\kappa a}{1+\kappa a} + \left(\frac{z}{\mathring{a}}\right)^2 \frac{36790}{\eta_0 DT}(\kappa a)^2\phi(\kappa a) \qquad (178)$$

where $\mathring{a}$ is the mean distance of approach of the ions in Ångström units. Another useful form for 1-1 type electrolytes is

$$\left(\frac{\mathfrak{M}}{c}\right)10^{20} = 1.0748\left(\frac{\lambda_1^0\lambda_2^0}{\Lambda^0}\right) - \frac{22.148}{\eta_0(DT)^{\frac{1}{2}}}\left(\frac{\lambda_1^0 - \lambda_2^0}{\Lambda^0}\right)^2 \frac{\sqrt{c}}{1+A'\sqrt{c}} + \frac{9.304 \times 10^7}{\eta_0(DT)^2}c\phi(A'\sqrt{c}) \qquad (179)$$

Here

$$\kappa a = A\sqrt{\Gamma} = A'\sqrt{c} \qquad (180)$$

$$A = \mathring{a}\frac{35.559}{(DT)^{\frac{1}{2}}} \qquad (181)$$

For convenience in utilizing these equations, values of $\phi(\kappa a)$ and $(\kappa a)^2\phi(\kappa a)$ as a function of κa are recorded in table 1.

The first term on the right of equation 179 is a constant and never greater than $\Lambda^0/4$. Since $(\lambda_1^0 - \lambda_2^0)$ is small for many salts, the second term is frequently negligible. As a glance at table 1 will show, $c\phi(A'\sqrt{c})$, which equals $(1/A')^2(\kappa a)^2\phi(\kappa a)$, is nearly constant at high concentrations and is never greater than 0.1. As a result, the variation of the diffusion coefficient is governed to a greater extent by the factor $[1 + c(\partial \ln y_\pm/\partial c)]$ than by $(\mathfrak{M}/c)$.

This term may be evaluated for 1-1 electrolytes from equation 129

$$\log y_\pm = \frac{-\mathfrak{S}_{(f)}\sqrt{c}}{1+A'\sqrt{c}} + 2Bc + D'c^2 - \log\left[\frac{d + 0.001c(2M_1 - M_2)}{d_0}\right] \qquad (182)$$

in which case

$$\left(1 + c\frac{\partial \ln y_\pm}{\partial c}\right) = 1 - \frac{1.1514\mathfrak{S}_{(f)}\sqrt{c}}{(1+A'\sqrt{c})^2} + 4.606Bc + 4.606D'c^2 - c\psi(d) \qquad (183)$$

where

$$\psi(d) \equiv \frac{\partial d/\partial c + 0.001(2M_1 - M_2)}{d + 0.001c(2M_1 - M_2)} \qquad (184)$$

Values of parameters of this equation for a number of important 1-1 electrolytes at 25°C. are given by Harned and Owen (41).

Observing that $c\phi(A'\sqrt{c})$ and its derivative with respect to c approach zero as c approaches zero, the limiting equation

$$\mathfrak{D} = \mathfrak{D}_0 - \mathfrak{S}_{(\mathfrak{D})}\sqrt{c} \qquad (185)$$

s obtained where

$$\mathfrak{D}_0 = 17.872 \times 10^{-10} T \left(\frac{\lambda_1^0 \lambda_2^0}{\Lambda^0} \right) \tag{185}$$

and

$$\mathcal{S}_{(\mathfrak{D})} = \frac{3.754 \times 10^{-3}}{D^{\frac{3}{2}} T^{\frac{1}{2}}} \left(\frac{\lambda_1^0 \lambda_2^0}{\Lambda^0} \right) + \frac{3.683 \times 10^{-8}}{\eta_0 D^{\frac{1}{2}} T^{-\frac{1}{2}}} \left(\frac{\lambda_1^0 - \lambda_2^0}{\Lambda^0} \right)^2 \tag{187}$$

The result obtained for $\mathfrak{D}_0$, the diffusion coefficient at infinite dilution, is of course the same as that obtained by Nernst and derived previously.

TABLE 1

Values of the functions $\phi(\kappa a)$ and $(\kappa a)^2\phi(\kappa a)$ for use in equations 178 and 179 (41)

κa	$\phi(\kappa a)$	$(\kappa a)^2\phi(\kappa a)$	κa	$\phi(\kappa a)$	$(\kappa a)^2\phi(\kappa a)$
0.000	∞	0	0.5	0.2651	0.06628
0.010	3.3550	0.00034	0.6	0.2054	0.07344
0.025	2.4695	0.00154	0.7	0.16304	0.07989
0.05	1.8273	0.00457	0.8	0.13194	0.08444
0.10	1.2342	0.01234	0.9	0.10844	0.08784
0.15	0.9245	0.02080	1.0	0.09033	0.09033
0.20	0.7277	0.02911	1.2	0.06477	0.09327
0.25	0.5907	0.03692	1.4	0.04813	0.09433
0.30	0.4899	0.04409	1.6	0.03676	0.09411
0.35	0.4130	0.05059	1.8	0.02876	0.09318
0.40	0.3527	0.05643	2.0	0.02292	0.09168
0.45	0.3044	0.06164	2.5	0.01391	0.08694
0.50	0.2651	0.06628	3.0	0.00908	0.08172

8. The dissipation function of a simple electrolyte for diffusion and conductance

In order to compute the dissipation function for the combined processes of electrical conductance and diffusion, we employ the electrochemical potential, $\bar{\mu}_i$, and the force, $\mathbf{k}_i$, obtained from its gradient. Thus

$$\bar{\mu}_i = \mu_i + e_i\psi \tag{188}$$

$$\mathbf{k}_i = - \nabla\bar{\mu}_i = - \nabla\mu_i + e_i\nabla\psi = - \nabla\mu_i + e_i\mathbf{X} \tag{189}$$

The equation relating the forces $\mathbf{k}_i$ to the flows $\mathbf{J}_k$ is

$$\mathbf{k}_i = \sum_{k=1}^{s} R_{ik}\mathbf{J}_k\,; \quad (i = 1, 2) \tag{190}$$

which corresponds to equation 97 and where

$$R_{ik} = R_{ki} \tag{191}$$

The dissipation function (equation 101) is given by

$$2F(\mathbf{J}, \mathbf{J}) = \sum_{i,k=1}^{s} R_{ik}\mathbf{J}_i\mathbf{J}_k = 2\phi(\mathbf{J}, \mathbf{J}) \tag{192}$$

and as a result of the above relations

$$\frac{\partial F}{\partial \mathbf{J}_i} = -\text{grad}\ \bar{\mu}_i = -\nabla\bar{\mu}_i = \mathbf{k}_i \tag{193}$$

For a simple electrolyte ($s = 2$),

$$\begin{aligned} \mathbf{k}_1 &= R_{11}\mathbf{J}_1 + R_{12}\mathbf{J}_2 \\ \mathbf{k}_2 &= R_{21}\mathbf{J}_1 + R_{22}\mathbf{J}_2 \\ 2F &= R_{11}\mathbf{J}_1^2 + 2R_{12}\mathbf{J}_1\mathbf{J}_2 + R_{22}\mathbf{J}_2^2 \end{aligned} \tag{194}$$

from which the operation given by equation 193 is seen to be valid.

In order to compute the dissipation function we first gather the detailed expression for the force $\mathbf{k}_j$. In the limit as the concentration approaches zero

$$\mathbf{k}_j = \rho_j\mathbf{v}_j = \mathbf{v}_j/\omega_j = \mathbf{J}_j/n_j\omega_j \tag{195}$$

At finite concentrations, owing to the time-of-relaxation effect and electrophoresis, $\Delta\mathbf{k}_j$,

$$\mathbf{k}_j = \rho_j\mathbf{v}_j - e_j\Delta\mathbf{X}_j - \rho_j\Delta\mathbf{v}_j \tag{196}$$

Upon introducing the detailed expression for ΔX_j from equation 139 and $\Delta\mathbf{v}_j$ given by equation 154 we obtain:

$$\begin{aligned} \mathbf{k}_j = \rho_j\mathbf{v}_j - e_j\left(\frac{e_1e_2\kappa}{3DkT}\cdot\frac{(1-\sqrt{q^*})n_2e_2}{(n_1e_1^2\omega_2 + n_2e_2^2\omega_1)}\right)(\omega_2\mathbf{k}_2 - \omega_1\mathbf{k}_1) \\ + \frac{2}{3\eta}\frac{\rho_je_je_\sigma n_\sigma\mathbf{k}_\sigma}{DkT\kappa(1+\kappa a)} - \frac{1}{3\eta}\left(\frac{\rho_j^{\frac{1}{2}}e_je_\sigma e^{\kappa a}}{DkT(1+\kappa a)}\right)^2 n_\sigma\mathbf{k}_\sigma\,Ei(\kappa a) \end{aligned} \tag{197}$$

Now

$$\begin{aligned} n_2e_2(\omega_2\mathbf{k}_2 - \omega_1\mathbf{k}_1) &= n_2e_2\omega_2\mathbf{k}_2 - n_1e_1\omega_1\mathbf{k}_1 \\ &= (e_2\mathbf{J}_2 + e_1\mathbf{J}_1) \end{aligned} \tag{198}$$

$$n_1\mathbf{k}_1 = \mathbf{J}_1\rho_1; \qquad n_2\mathbf{k}_2 = \mathbf{J}_2\rho_2 \tag{199}$$

$$\phi(\kappa a) = \frac{e^{2\kappa a}}{(1+\kappa a)^2}Ei(\kappa a) \tag{200}$$

and

$$\omega_1 = \frac{1}{\rho_1}; \qquad \omega_2 = \frac{1}{\rho_2} \tag{201}$$

Utilizing these relations, we obtain for a 1-1 electrolyte

$$\begin{aligned} \mathbf{k} = \mathbf{k}_1 + \mathbf{k}_2 = \mathbf{J}_1\rho_1/n_1 + \mathbf{J}_2\rho_2/n_2 \\ - \frac{e_1e_2\kappa}{3DkT}\frac{\rho_1\rho_2(1-\sqrt{q^*})}{(n_1e_1^2\rho_1 + n_2e_2^2\rho_2)}(e_1+e_2)(e_1\mathbf{J}_1 + e_2\mathbf{J}_2) \\ + \frac{2}{3\eta DkT\kappa(1+\kappa a)}(\rho_1e_1 + \rho_2e_2)(e_1\rho_1\mathbf{J}_1 + e_2\rho_2\mathbf{J}_2) \\ - \frac{\phi(\kappa a)}{3\eta(DkT)^2}(\rho_1e_1^2 + \rho_2e_2^2)(e_1^2\rho_1\mathbf{J}_1 + e_2^2\rho_2\mathbf{J}_2) \end{aligned} \tag{202}$$

The dissipation function takes the form

$$\begin{aligned} 2F &= \mathbf{R}_{11}\mathbf{J}_1^2 + 2R_{12}\mathbf{J}_1\mathbf{J}_2 + R_{22}\mathbf{J}_2^2 \\ &= (\rho_1/n_1)\mathbf{J}_1^2 + (\rho_2/n_2)\mathbf{J}_2^2 - \frac{e_1 e_2 \kappa}{3DkT}\frac{\rho_1\rho_2(1-\sqrt{q^*})}{(n_1 e_1^2 \rho_1 + n_2 e_2^2 \rho_2)}[e_1\mathbf{J}_1 + e_2\mathbf{J}_2]^2 \\ &\quad + \frac{2}{3\eta DkT(1+\kappa a)\kappa}[e_1\rho_1\mathbf{J}_1 + e_2\rho_2\mathbf{J}_2]^2 - \frac{\phi(\kappa a)}{3\eta(DkT)^2}[e_1^2\rho_1\mathbf{J}_1 + e_2^2\rho_2\mathbf{J}_2]^2 \end{aligned} \tag{203}$$

The correctness of this expression may be proved by differentiation with respect to $\mathbf{J}_1$ and $\mathbf{J}_2$ according to equation 193 and comparing the result with the preceding expression for **k**.

The first two terms of this equation represent the dissipation function when the effects of Coulombic forces are neglected. The second is the time-of-relaxation effect which, as previously shown, vanishes in the case of diffusion alone when both ions migrate with the same velocity. The last two terms represent the electrophoretic effect.

For diffusion

$$\mathbf{v} = \mathbf{v}_1 = \mathbf{v}_2 = \mathbf{J}_1/n_1 = \mathbf{J}_2/n_2 \tag{204}$$

and

$$2F = \left[n_1\rho_1 + n_2\rho_2 + \frac{2e_1^2 n_1^2}{3\eta DkT}\frac{(\rho_1-\rho_2)^2}{(1+\kappa a)\kappa} - \frac{\phi(\kappa a)(e_1 n_1)^2(\rho_1 e_1 - \rho_2 e_2)^2}{3\eta(DkT)^2}\right]\mathbf{v}^2 \tag{205}$$

We note that for an electrolyte dissociating into two ions, the electrophoretic terms will vanish if the frictional coefficients of the ions are identical.

V. EXPERIMENTAL METHODS FOR THE MEASUREMENT OF DIFFUSION COEFFICIENTS[6]

A. BOUNDARY CONDITIONS AND ADAPTATIONS OF FICK'S LAWS FOR THE MEASUREMENT OF DIFFUSION COEFFICIENTS

The experimental methods for measuring diffusion coefficients may be arranged under three classes, which depend on whether or not a steady state is reached or whether the entire solution is maintained in a closed system. The application of Fick's laws then depends on the experimentally imposed boundary conditions.

1. *Free diffusion*

Free diffusion refers to diffusion from an initially sharp boundary between two solutions at different concentrations, or between solution and solvent, in a vertical apparatus in which the composition of the liquid media at the top and bottom remain constant. In order to adapt equations 16 or 17 for practical calculations, use is made of the fact, first observed by Boltzmann (5), that the variables x and t always occur in the ratio x^2/t. Consequently, since $x = 0$ when $t = 0$, we may let $y = x/t^{\frac{1}{2}}$ and introduce this new variable in the differential equations. Thus, equation 16 becomes

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[6] For this part of the review, the author is particularly indebted to Dr. Lewis G. Longsworth of the Rockefeller Institute for Medical Research, whose paper (61) on the theory of measurement of diffusion coefficients has been used as a basis for this discussion and who has kindly permitted the reproduction of some of the figures employed in describing the optical methods.

$$\frac{y}{2}\frac{\mathrm{d}n}{\mathrm{d}y} = -\frac{\mathrm{d}}{\mathrm{d}y}\mathfrak{D}\frac{\mathrm{d}n}{\mathrm{d}y} \tag{206}$$

and

$$\begin{aligned} -\mathfrak{D} &= \frac{\mathrm{d}y}{\mathrm{d}n}\int_0^n \frac{y}{2}\,\mathrm{d}n = \frac{1}{2t}\frac{\mathrm{d}x}{\mathrm{d}n}\int_0^n x\,\mathrm{d}n \\ &= \frac{1}{2t}\frac{\mathrm{d}x}{\mathrm{d}n}\int_{-\infty}^{x} x\frac{\mathrm{d}n}{\mathrm{d}x}\,\mathrm{d}x \end{aligned} \tag{207}$$

This equation has been employed as a basis for computing $\mathfrak{D}$ by the float method (29). Also, since $\mathrm{d}n/\mathrm{d}x$ is usually very nearly proportional to the refractive-index gradient, it is the basic relation for the development of the optical methods of evaluating $\mathfrak{D}$. If $\mathfrak{D}$ is assumed to be independent of y, equation 206 becomes:

$$\frac{y}{2}\frac{\mathrm{d}n}{\mathrm{d}y} = -\mathfrak{D}\frac{\mathrm{d}^2n}{\mathrm{d}y^2} \tag{208}$$

Upon integration and resubstitution of x and t,

$$\frac{\mathrm{d}n}{\mathrm{d}x} = \frac{1}{It^{\frac{1}{2}}}e^{-x^2/4\mathfrak{D}t} \tag{209}$$

where I is a constant of integration. Now

$$\int_{-\infty}^{\infty}\frac{\mathrm{d}n}{\mathrm{d}x}\,\mathrm{d}x = n_1 - n_0 \tag{210}$$

which is the total area under the curve of $\mathrm{d}n/\mathrm{d}x$ *versus* x. Therefore the concentration difference of the two ends of the diffusion column is given by

$$\begin{aligned} n_1 - n_0 &= \frac{1}{It^{\frac{1}{2}}}\int_{-\infty}^{\infty} e^{-x^2/4\mathfrak{D}t}\,\mathrm{d}x \\ &= \frac{2}{It^{\frac{1}{2}}}\int_0^{\infty} e^{-x^2/4\mathfrak{D}t}\,\mathrm{d}x \\ &= \frac{2}{It^{\frac{1}{2}}}\sqrt{\pi\mathfrak{D}t} \end{aligned} \tag{211}$$

and

$$I = \frac{2\sqrt{\pi\mathfrak{D}}}{n_1 - n_0} \tag{212}$$

Upon eliminating I from equation 209 we obtain Wiener's equation (105)

$$\frac{\mathrm{d}n}{\mathrm{d}x} = \frac{(n_1 - n_0)}{2\sqrt{\pi\mathfrak{D}t}}e^{-x^2/4\mathfrak{D}t} \tag{213}$$

from which

$$n = \frac{(n_1 - n_0)}{2\sqrt{\pi\mathfrak{D}t}} \int_{\infty}^{x} e^{-x^2/4\mathfrak{D}t}\, dx \tag{214}$$

These equations, which give the distribution of solute as a function of height, x, find their principal application in the use of optical methods for the determination of diffusion coefficients of macromolecules. Although it was recognized that the gradient of the refractive index could be used to measure a concentration gradient (33), the first optical system used quantitatively for this purpose was constructed by Wiener (105). Important technical advances were made some years later by Thovert (96, 97) and others (45, 83, 104), who used optical methods for measuring diffusion coefficients. The three modern methods, the absorption, the "schlieren", and the scale, have been developed to a high degree of precision at the Chemical Institute of Upsala by Svedberg (93), Tiselius (98), and Lamm (37, 52–56). The investigations of Longsworth (60–66), who is responsible for the "schlieren scanning" method, are the last word in precision technique by optical methods of this nature.

These experimental procedures have found their principal application in determining diffusion coefficients of macromolecules and in electrophoretic analysis. With the exception of Clack's researches (9, 10), very few accurate diffusion coefficients of electrolytes have been determined by optical means. On the other hand, these procedures can be carried out in a short time as compared with that usually required for the study of the slow process of diffusion and for this reason may prove important in future studies.

In addition to the optical methods, the microindicator method (25–28, 84, 99, 108) and the float method of Gerlach (29, 106) have been employed for the determination of diffusion coefficients from systems undergoing free diffusion. These have not yielded nor do they promise to yield as accurate results as other methods and will not be mentioned further.

2. *Steady-state diffusion*

The diffusion upwards in a vertical column from a reservoir at the bottom containing a solution at one concentration into a reservoir at the top containing solution at another lower concentration, or into pure solvent, ultimately attains a steady state in which, at each height in the column, the flow J is constant. With such a column Clack (9, 10), with an accurate series of measurements of some salt solutions in water, was able to determine both the flow J and the concentration gradient dn/dx. From these, he was able to compute the differential diffusion coefficient by Fick's first law:

$$\mathfrak{D} = -J(\mathrm{d}n/\mathrm{d}x)^{-1} \tag{215}$$

These results are among the very few direct determinations of differential diffusion coefficients and their accuracy at concentrations of 0.05 N and higher are estimated to be within 1.5 per cent. These results constitute one of the few important contributions to our knowledge of electrolytic diffusion but were not obtained at sufficiently low concentrations to afford an accurate test of theory. Clack (9) also used a gravitational method in a system in a steady state of diffusion to determine integral diffusion coefficients.

The diaphragm method of McBain (67, 68, 101) and Northrup (74), which permits a solute on one side of a sintered-glass membrane to diffuse into water or a more dilute solution on the other side of the membrane, approaches the steady condition if the solutions employed are sufficiently large. Under this condition, the diaphragm method can be treated as a steady-state problem, since the concentrations on the two sides of the membrane will not alter appreciably. This method has been shown by McBain (67) and Dawson (16), and particularly by Hartley and Runnicles (43) and Gordon (31, 32), to yield accurate relative results.

3. Restricted diffusion

When diffusion from an initially sharp boundary takes place in a solution which completely fills a vessel and is permitted to proceed for a sufficient length of time, the concentrations at all heights in the apparatus will change continually. In a vertical column of height a, the flow J will be zero when $x = 0$ and $x = a$, since the solute is confined within the vessel. In this case, the differential equation

$$\frac{\partial n}{\partial t} = -\mathfrak{D} \frac{\partial^2 n}{\partial t^2} \tag{216}$$

with boundary conditions

$$\begin{aligned} \frac{\partial n}{\partial x} &= 0 \quad \text{for} \quad \begin{cases} x = 0 \\ x = a \end{cases} \\ n &= f(x) \quad \text{for} \quad t = 0 \\ n &\neq \infty \quad \text{for} \quad t = \infty \end{aligned} \tag{217}$$

may be satisfied by a solution arranged in a Fourier series. Thus

$$n = n_0 + \sum_{m=1}^{\infty} A_m e^{-m^2\pi^2\mathfrak{D}t/a^2} \cos \frac{m\pi x}{a} \tag{218}$$

where

$$n_0 = \frac{1}{a}\int_0^a f(x)\, dx; \qquad A_m = \frac{2}{a}\int_0^a f(x) \cos \frac{m\pi x}{a}\, dx \tag{218a}$$

Obviously, $n = n_0$ when $t = \infty$, in agreement with the condition that no solute can enter or leave the system. As an example of restricted diffusion, we cite the layer analysis method, used by many investigators from the very beginning of the study of diffusion and culminating in the precise study of the diffusion of potassium chloride in water by Cohen and Bruins (12). Among these investigations we mention particularly the extensive series of results of Öholm (75, 76). Unfortunately, determinations like those of Cohen and Bruins yield integral diffusion coefficients and, as pointed out by Gordon (32), it is difficult to compare such results with the integral values obtained by the diaphragm-cell method.

Electromotive-force measurements (82, 85, 89, 103) and conductance measurements (44, 56, 73) have also been applied to measure diffusion coefficients in vessels in which restricted diffusion occurred. We shall describe in some detail a method which involves the measurement of electrical resistance at the top and bottom of a carefully designed vertical cell. By such a technique (39, 40), the diffusion coefficient of potassium chloride in water from 0.001 N to 0.01 N has been determined with an accuracy of the order of 0.1 per cent. These results serve for the first time to prove the validity of the limiting equation of Nernst (71) and to test the theory of Onsager and Fuoss (80).

B. FREE DIFFUSION. A GENERAL SURVEY OF OPTICAL SYSTEMS USED FOR DETERMINATION OF DIFFUSION COEFFICIENTS[7]

A general idea of how a concentration gradient in a solution deflects light from the normal path is clearly illustrated by figure 1. A horizontal slit to the left of lens L is focussed by this lens in the plane P. If the cell G contains a homogeneous liquid or solution, the wave front may be represented by the circular arc *ab* and the image of the slit will be focussed at x_0. Now suppose that free diffusion takes place at a sharp boundary between solution at the bottom and solvent at the height h_0. After a period of time, concentration gradients of magnitudes indicated by the intensity of the shading will be present and the new wave front will have the form ah_0b'. Since the refractive index of the more concentrated solution is the greater, the velocity of light will be less in the lower portion and $h''b'$ will be retarded relative to ah'. However, the light passing through these regions will still converge at x_0. On the other hand, the light which passes through the portion $h'h_0h''$ will converge at x_1, as indicated by the lines drawn normal to the wave front, and if the diffusion is not affected by turbulence or other disturbing factors will form a series of light and dark bands,

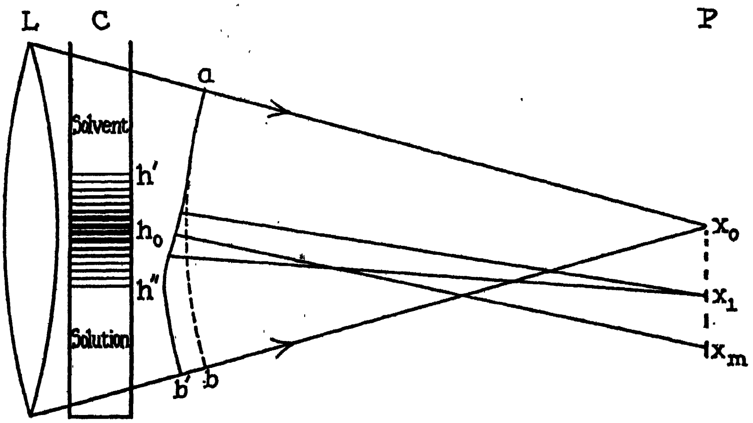

FIG. 1. Diagram illustrating the interference phenomena accompanying the deflection of light by gradients of refractive index in a freely moving boundary (Longsworth).

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[7] See reference 1, which contains a review of optical methods.

which depend on the phase relations of the light. These bands increase in width from x_0 to x_m, at which level a diffuse band appears corresponding to the normal of the inflection point of the wave front.

In figure 2 a photograph of such a system of bands taken by Longsworth (61) is reproduced. He suggested that a quantitative theory of the distribution of these bands for the case of ideal free diffusion would be of considerable value for the application of optical methods.[8]

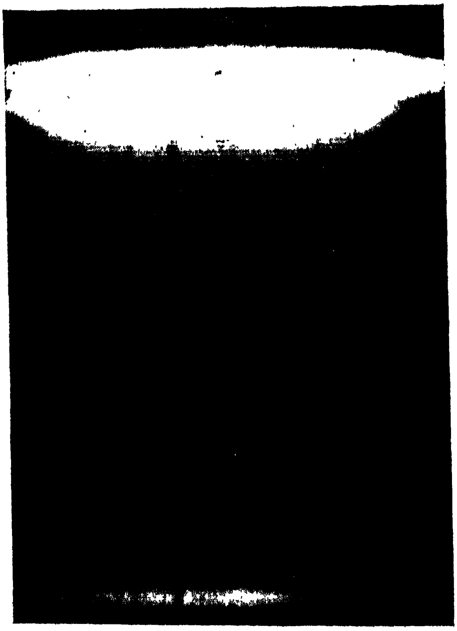

FIG. 2. Photograph at the schlieren diaphragm (Longsworth)

1. *The "schlieren" method for the observation of boundaries and bands*

The adaptation of Toepler's schlieren (shadow) method by Tiselius (98) for the observation of boundaries formed during the electrophoresis of macromolecules inaugurated an important development in the modern optical technique of measuring concentration gradients. By this method, a graph of a gradient of the refraction index dn/dx *versus* the vertical distance x in the cell may be constructed.[9] The method of Tiselius has been greatly improved in accuracy and ease of manipulation by Longsworth (60–66), who devised the

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[8] In a private communication, Dr. Longsworth has informed me that as a result of this suggestion, Dr. Gerson Kegeles at the University of Wisconsin has developed a theory of this phenomenon by means of which the diffusion coefficient for ideal diffusion may be computed from the spacing of the bands.

[9] In this section n has been used to denote refractive index, in conformity with common usage.

"schlieren scanning" method by which a photograph of the plot of dn/dx against x can be obtained directly.[10]

A simple diagram of the optical system is shown in figure 3. An image of a horizontal slit is brought to focus in the plane at P by the lens D. At P an opaque diaphragm with a sharp horizontal upper edge is placed. This can be moved vertically. We note that the schlieren diaphragm is at the same position as the plane P in figure 1, so that the photograph (figure 2) was taken at this position. The objective lens O serves to focus the image of the cell E upon the photographic plate G.

If gradients of the refractive index are present in the cell, the light through these will be deflected downward as shown in figure 1. If the diaphragm is placed so as to intercept these deflected pencils, dark horizontal bands appear at the photographic plate. These shadows are known as the schlieren bands, which with electrophoretic boundaries can be made sharp by suitable lenses and a sufficiently narrow slit S. The displacement of the diaphragm from the position of the undeviated to the deviated image is proportional to the gradients of the refractive index in the cell conjugate to the edges of the bands.

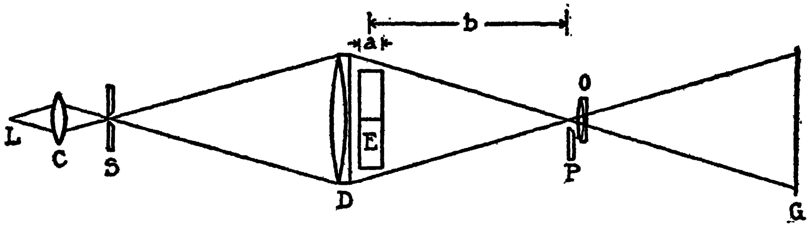

FIG. 3. Schematic diagram of the optical system employed in the schlieren method (Longsworth).

The facility with which the schlieren method can be applied has been greatly increased by Longsworth. A vertical slit is placed at the photographic plate and the latter is moved horizontally past the slit. Simultaneously, the schlieren diagram is moved vertically until the position of the undeviated slit image is reached. By this method a transparent area caused by the interception of the deviated pencils of light appears on the plate. The contour line of this area is a graph of the gradient dn/dx against the vertical position x.

The theory of the deflection of light by an inhomogeneous liquid system shows that if δ is the displacement of the schlieren diaphragm

$$\delta = ab\,\mathrm{d}n/\mathrm{d}x \qquad (219)$$

where a is thickness of the liquid in the cell and b is the optical distances indicated in figure 3. Lamm's (54) recent contribution to this subject contains detailed considerations of the optics of the scale and schlieren methods with particular emphasis upon the scale method. The theory of ray deflection is developed exhaustively.

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[10] Special attention is suggested to the article by Longsworth (64), which is a masterpiece of description of this experimental method.

Now the integral which equals the area under the plot derived from the schlieren scanning photographs is

$$\int \delta \, \mathrm{d}x = \int ab \frac{\mathrm{d}n}{\mathrm{d}x} \, \mathrm{d}x = ab\Delta n \tag{220}$$

This equation has proved of importance in the study of electrophoresis, where a substance is concentrated in a thin layer. For this case, the increment in refractive index, Δn, is nearly proportional to the concentration of substance. Consequently the schlieren scanning method can be used conveniently to estimate the total concentration of the substance and the distribution of concentration through the boundary.

2. The scale method of Lamm (54)

The optical system of the scale method consists of a lens by means of which a scale is focussed on a plane photographic plate. A vertical cell with parallel sides is placed between the scale and the lens at a distance b from the scale. The scale is first photographed through a distance a of a homogeneous solvent or solution. It is photographed again later through the inhomogeneous media produced by the diffusing substances or electrophoretic boundary. The refractive-index gradients cause deviations in the images of the lines of the scale on the photographic plate which can be determined by comparison with the normal scale-line positions. For this purpose a comparator is used. In this optical system the deviated line is somewhat out of focus, but this condition can be remedied to a large extent by employing a lens with a long focal length.

Lamm (54) has made a very careful and exhaustive theoretical and experimental investigation of the scale method, and his dissertation should be read by all those seriously interested in the subject. Geometric optics shows that the displacement of a line from its normal position on the photographic plate is given by the equation

$$Z = Gab \, \mathrm{d}n/\mathrm{d}x \tag{221}$$

where a and b are the distances mentioned above and G is a magnification factor (photographic enlargement).

The photographic deviation Z and scale reading z are arbitrary but depend on the refractive-index gradient $\mathrm{d}n/\mathrm{d}x$, which we shall denote by X, and the x-dimension in the cell. The projection factor is $\mathrm{d}x/\mathrm{d}z$ or F, which when introduced in equation 22 yields

$$\mathrm{d}n = \frac{F}{Gab} Z \, \mathrm{d}z \tag{222}$$

and the integral

$$n - n_0 = \frac{F}{Gab} \int_{z_{n_0}}^{z_n} Z \, \mathrm{d}z \tag{223}$$

In order to determine the diffusion coefficient from these measurements of free diffusion, we resort to equation 213, which was derived by Wiener (105).

Thus

$$X = \mathrm{d}n/\mathrm{d}x = \frac{n_1 - n_0}{2\sqrt{\pi \mathfrak{D} t}}\, e^{-x^2/4\mathfrak{D}t} = \frac{\Phi}{2\sqrt{\pi \mathfrak{D} t}}\, e^{-x^2/4\mathfrak{D}t} \tag{224}$$

where Φ is the area under the (X, x) graph. This relation is recognized to have the form of the error or probability law, and its general validity in the case of pure free diffusion is well illustrated by the diagram of Wiener's optical system in figure 4. The illuminated slit at an angle of 45° is focussed upon a screen (G) by a lens of large focal length. If the solution in the cell (C) is homogeneous, the image of the slit is the straight dashed line. If a diffusion gradient is produced from an originally sharp boundary at h_0 between a solution at the bottom and solvent at the top, the light rays are bent downward and the image upon G has the characteristic form of the error function, *ab′c*. Somewhat later Thovert, by interposing a cylindrical lens between the cell and a photograph, obtained a horizontal image. One of his photographs of an experiment in which 0.85 *N* sodium chloride is diffusing into water is shown in figure 5. The three curves correspond to different times during the diffusion process.

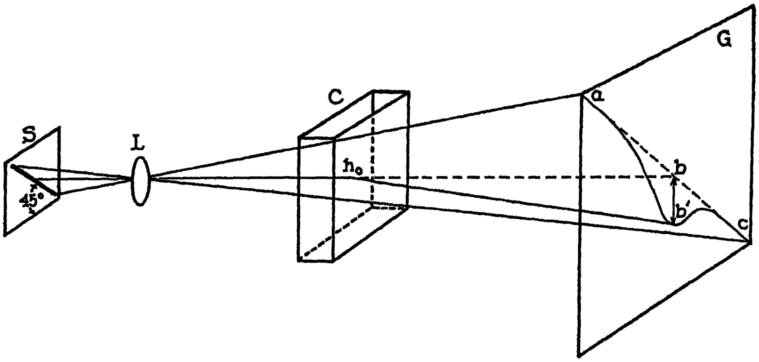

FIG. 4. The optical system of Wiener (Longsworth)

The actual experimental observations in the scale method are deviations of the scale lines on the photograph, Z, and the distances of the scale lines, z, from some arbitrary origin. Since these results depend on equation 224, which has the form of a Gauss distribution function, they may be computed by the methods of mathematical statistics.[11] To this end, the experimental values of the ordinates Z are plotted against the abscissae z. The z-axis is then divided into a large number of equal parts of length w, and the scale values corresponding to each of these divisions are obtained from the ordinates of the graph. From an arbitrary origin ($s = 0$) the divisions are numbered 1, 2, etc. to the right and −1, —2, etc. to the left of this arbitrary origin.

The ideal frequency curve $S = f(s)$ of Charlier (11)

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[11] K. Pearson (Trans. Roy. Soc. (London) **A185**, 71 (1894)). For a simple textbook treatment see B. H. Camp, *The Mathematical Part of Elementary Statistics*, Heath and Company, New York (1934).

$$S = \frac{N\omega}{\sigma\sqrt{2\pi}} e^{-(s-\beta)^2\omega^2/2\sigma^2} \tag{225}$$

is suitable for this computation. In this relation, S is the frequency of a given statistical element, N is the total number of statistical elements, s is the class number, ω is the class breadth, σ is the standard deviation, and β is the distance of the arbitrary origin from the centroid.

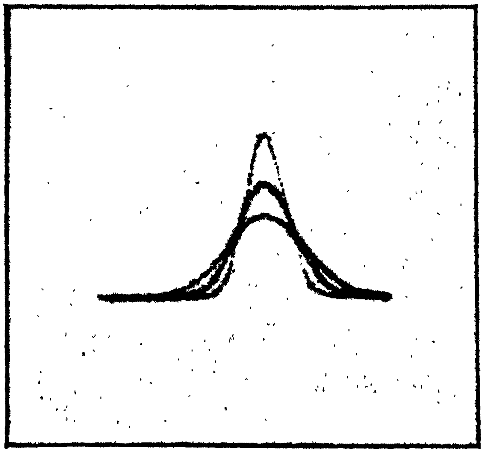

FIG. 5. Photograph obtained by Thovert of 0.85 N sodium chloride diffusing into water (Longsworth).

This equation becomes analogous to equation 213

$$X = \frac{\mathrm{d}n}{\mathrm{d}x} = \frac{\Phi}{2\sqrt{\pi\mathfrak{D}t}} e^{-x^2/4\mathfrak{D}t} \tag{226}$$

upon substituting

$$x = (s - \beta)\omega \tag{227}$$

$$\Phi = N\omega = \text{area} \tag{228}$$

$$\sigma^2 = 2\mathfrak{D}t \tag{228}$$

The quantities N, β, and σ/ω are determined from the theory by the relations

$$N = \Sigma S \tag{230}$$

$$\beta = \frac{\Sigma Ss}{\Sigma S} \tag{231}$$

and

$$\left(\frac{\sigma}{\omega}\right)^2 = \frac{\Sigma s^2 S}{\Sigma S} - \left(\frac{\Sigma Ss}{\Sigma S}\right)^2 \tag{232}$$

The latter equation is derived from consideration of the area and the second moment. The maximum value of the frequency S_m occurs when $(S - \beta)\omega$ is zero, whence

$$S_m = \frac{N\omega}{\sigma\sqrt{2\pi}} = \frac{(\Sigma S)\omega}{\sigma\sqrt{2\pi}} \tag{233}$$

These later equations yield

$$\mathfrak{D} = \left[\frac{\Sigma s^2 S}{\Sigma S} - \left(\frac{\Sigma S s}{\Sigma S}\right)^2\right]\frac{\omega^2}{2t} \tag{234}$$

and

$$\mathfrak{D} = \frac{(\Sigma S)^2}{S_m^2}\,\frac{\omega^2}{4\pi t} \tag{235}$$

which have been frequently employed for the evaluation of the diffusion coefficient. Since the calculations are made from the Z, z graph, the projection constants of the apparatus must be employed for determining ω and N. In order to obtain the mean position of the boundary or to change the origin to the centroid, $\beta\omega$ is added to the s readings.

Further, in order to compare the frequency curve with the corresponding normal ideal Gauss curve, let

$$\xi = \frac{(s - \beta)\omega}{2\sigma} \tag{236}$$

and

$$\Xi = \frac{5\sigma}{N\omega}S \tag{237}$$

whence

$$\Xi = \frac{5}{\sqrt{2\pi}}e^{-2\xi^2} \tag{238}$$

The numbers 5 and 2 are arbitrary and are chosen for convenience in drawing the graph.

The free diffusion of a pure solute into a solvent has been shown in many instances to conform to the theory. Consequently, agreement with theory has been interpreted as proving that the solute is a homogeneous substance and that deviations from the ideal curve indicate that the diffusing substance is a mixture. The principal application of this latter reasoning has found the basis for numerous studios of diffusion of polydisperse systems containing macromolecules (3, 4).

That the real diffusion graph conforms to the ideal normal Gauss distribution is illustrated by figure 6. These are plots of Lamm's results for the diffusion of 0.1 N potassium chloride into water at 20°C. The bottom curve is the plot of the photographic displacements Z *versus* the distances z on the arbitrary scale with origin corresponding to the maximum of the graph. The top curve shows the same results in normal coördinates, Ξ and ξ. The agreement between the experimental and theoretical results is good.

Lamm (54) computed the diffusion coefficient for this case by a number of methods involving the mathematical properties of these curves and obtained values of 1.44 to 1.47 cm.^{2} per day or 1.67 to 1.70 $\times$ 10^{-5} cm.^{2} per second. This agrees closely with the result, 1.448 cm.^{2} per day, obtained by Cohen and Bruins (12) by the layer analysis method. However, this agreement must be taken cautiously, since it is difficult to interpret exactly results obtained by the layer analysis method (see Section V, D, 1).

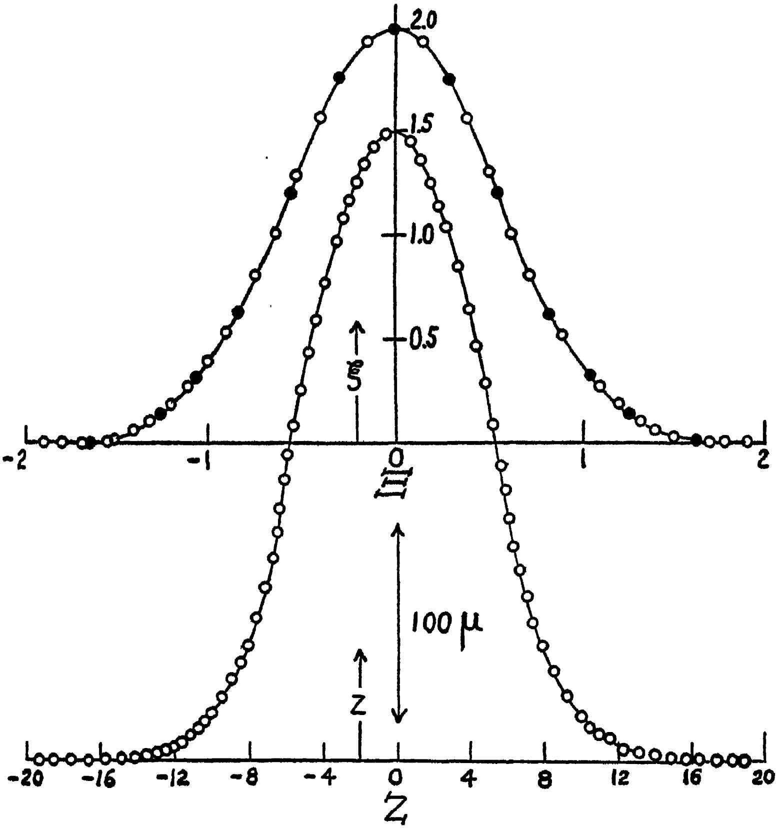

Fig. 6. Z,z and Ξ,ξ plots employed for the calculation of the diffusion coefficient by the scale method.

C. STEADY-STATE DIFFUSION

1. The measurement of the differential diffusion coefficient according to Clack (9, 10)

The experiments of Clack which led to the determination of the differential diffusion coefficients of potassium and sodium chlorides and potassium nitrate from 0.05 N to high concentrations are one of the most important contributions to this field. Clack's apparatus is represented in figure 7. A salt from a saturated solution is allowed to diffuse upward into water in the cell C until a steady state of diffusion is reached. The refractive index and its gradient are measured by an optical system consisting of a horizontal slit S focussed by the lenses L_1 and L_2 upon an ocular E. A screen B with a compound slit Y is interposed between the lens L_1 and the cell. This screen can be moved vertically along the column. Clack found that the interference pattern formed by the double slit in the ocular yielded a sharply defined central band. The displacement of the pencil of rays produced by the refractive-index gradient in the cell from its normal position can be measured by moving the ocular in a vertical direction.

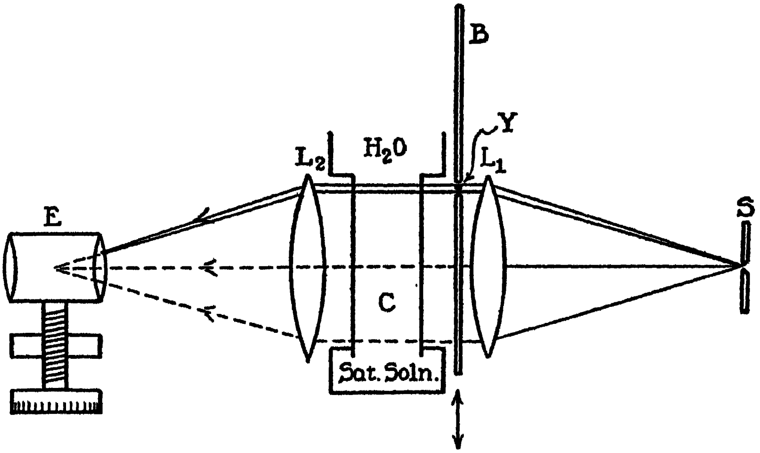

FIG. 7. The optical system used by Clack for the determination of differential diffusion coefficients (Longsworth).

As in the previous optical methods, this displacement δ is given by the equation

$$\delta = ab\, dn/dx \tag{239}$$

where a is the distance through which the light passes in the liquid, b is the distance from C to E, and x is position in the cell at the height of the slit Y. Knowing these dimensions and δ, dn/dx at any height z may be determined, and if n is known as a function of the concentration c, both c and dc/dx may be evaluated at the position x.

In order to evaluate the differential diffusion coefficient by this method, it is necessary to determine the solute flow J, and also to make a correction for the solvent counterflow. With this in mind, we write Fick's first law

$$\mathfrak{D} = -J_1\phi(dc/dx)^{-1} \tag{240}$$

where ϕ is the frame-of-reference factor. In Clack's final experiments, the lower reservoir of the cell contained solid salt in equilibrium with its saturated solution. When the steady state is attained, the flow J_1 is a constant and the lower reservoir loses J_1 grams of salt and gains J_0 grams of water per unit cross section in unit time. The quantities J_1 and J_0 can be computed from the weight of salt dissolved, the density of the salt, and the concentration and density of the saturated solution.

We have pointed out that the flow has usually been defined relative to a fixed plane of reference and have indicated by equation 2, which defines a bulk velocity, how the flow can be defined relative to a frame of reference moving with the solvent. It is clear that in dilute solutions when extremely small volume changes are involved there will be no appreciable difference between results obtained upon these two frames of reference. On the other hand, in Clack's cell the concentration of salt through the column varies from that of the saturated solution to zero, so that the results obtained at a reference fixed with respect to the apparatus must be corrected for solvent counterflow and volume changes. Clack derived expressions for this correction.

2. The diaphragm-cell method

The use of sintered-glass diaphragm cells for measuring diffusion coefficients has been shown by McBain (67) and especially by Gordon (14, 32) and Hartley and Runnicles (43) to be convenient and accurate. Since the diffusion process is slow, measurements in vessels without diaphragms are most sensitive to any factors which, like local gradients of temperatures or initial stirring, may cause convection. In the diaphragm cell, the diffusion is confined to the pores of a sintered-glass or an alundum diaphragm, so that errors caused by agitation by thermal or mechanical means are reduced to a minimum.

The apparatus used by McBain and Liu (68) and Gordon (14, 32) consisted of a bell-shaped vessel with a stopcock at the top and a flat diaphragm of sintered glass at the bottom. This (inner) vessel is filled with the more concentrated solution and is suspended in such a way that the diaphragm just touches the surface of a weaker (outer) solution. The diaphragm must be adjusted horizontally and must be in contact with the outer solution over its entire surface. The whole apparatus is placed in an air thermostat. In this form of apparatus only density stirring can take place.

Mouquin and Cathcart (69) have described a glass cell containing a sintered-glass diaphragm at the middle, which was stirred by rotating end over end. Hartley and Runnicles (43) obtained stirring by rotation of their cell about its axis at an angle inclined to the vertical. The upper solution was stirred by a glass sphere resting on the diaphragm, and the lower solution by a lighter glass sphere which pressed against the bottom surface of the diaphragm. The concentration in the inner compartment of the apparatus was measured by conductance.

The obvious disadvantage of the diaphragm-cell method consists in the fact that each cell possesses a cell constant which must be evaluated by employing a solution containing a solute with a known diffusion coefficient. Up to very recently, no determination of a diffusion coefficient of an electrolyte could be relied on to within a few per cent, so that no really satisfactory cell-constant calibration could be made.

On the other hand, by careful experimentation resulting from critical analysis of the sources of error, it has been shown that diaphragm-cell measurements are capable of high accuracy and reproducibility. The theory of the method, originally developed by Northrup and Anson (74) and McBain and Liu (68), is very simple. The diaphragm is assumed to be made up of a large number of parallel pores of a length l and of an effective cross-sectional area, A. The solutions on both sides of the diaphragm are assumed to have uniform composition and the transfer of matter from the inner to the outer solution occurs only by the process of diffusion. The membrane must be such that any stream-lined flow is prohibited. If c' and c'' change slowly, the solute distribution throughout the diaphragm will approach that of a steady-state process. The concentration gradient across the diaphragm if $\mathfrak{D}$ is independent of the concentration will be $(c' - c'')$ and will be constant throughout the diaphragm. If the quantity, dq, of solute diffuses in the time dt, then

$$\mathrm{d}q = \frac{\mathfrak{D}A}{l}(c' - c'')\,\mathrm{d}t \tag{241}$$

which is the differential form of equation for the diaphragm cell. The concentrations c', c'' and the quantity $\mathrm{d}q/\mathrm{d}t$ may be measured, but A/l is not known and must be determined from the known diffusion coefficient of some solute. Northrup and Anson used this differential form of equation. There are obvious practical advantages of extending the measurements over longer time intervals and most of these measurements have employed the result obtained by integration. The change in the quantity of solute in one compartment in a time dt must equal the amount which has diffused in or out of this compartment. If V' and V'' are the volumes of the inner and outer solutions, respectively, then

$$\begin{aligned} V'\mathrm{d}c' + (\mathfrak{D}A/l)(c' - c'')\,\mathrm{d}t &= 0 \\ V''\mathrm{d}c'' + (\mathfrak{D}A/l)(c'' - c')\,\mathrm{d}t &= 0 \end{aligned} \tag{242}$$

where dc' and dc'' are the concentration changes. Upon subtracting the second of these equations from the first and writing Δc for $(c' - c'')$, we obtain:

$$\frac{\mathrm{d}(\Delta c)}{\Delta c} + \beta\mathfrak{D}\,\mathrm{d}t = 0 \tag{243}$$

where the cell factor constant β is given by

$$\beta = \frac{A}{l}\left(\frac{1}{V'} + \frac{1}{V''}\right) \tag{243a}$$

Assuming that $\mathfrak{D}$ is not a function of c, and that the cell remained constant during an experiment, equation 243 can be integrated from the initial concentrations c_0', c_0'' to the final concentration c_f', c_f'' to yield

$$\ln\frac{\Delta c_f}{\Delta c_0} = -\beta\mathfrak{D}t \tag{244}$$

the equation most frequently used in determining the diffusion coefficient by this method.

An analytical and critical survey of the diaphragm-cell technique has been made by Gordon (32). In addition to the discussion of the sources of errors and their elimination, he has shown that the density-stirred method yields results which agree with those obtained by mechanical stirring. He also shows that the diaphragm-cell method may be safely treated as a steady-state process. We shall reserve for a later section the discussion of the selection of a standard diffusion coefficient required for the calibration of the cell.

Gordon (32) has made an important contribution in showing how the differential diffusion coefficient can be evaluated from diaphragm-cell measurements. Let us assume that $\mathfrak{D}$ is a function of the concentration so that

$$\mathfrak{D}/\mathfrak{D}_0 = 1 + f(c) \tag{245}$$

where $\mathfrak{D}_0$ is the value of $\mathfrak{D}$ at infinite dilution. If this is the case, then the derivation of equation 244 is incorrect. Now for convenience, an integral diffusion coefficient $\mathfrak{D}_i$ may be defined by equation 244, since ΔC_f, ΔC_0, and β are known. By the following procedure, one may find the concentration c_i at which $\mathfrak{D}_i$ equals the differential diffusion coefficient $\mathfrak{D}$. For a steady-state process at a given time, the flow $\mathfrak{D}\partial c/\partial x$ is constant from the top ($x = 0$) to the bottom ($x = l$) of the membrane, for if this were not the case the transport of matter would be unequal at different areas on the diaphragm. At this instant, let the inner and outer concentrations be c' and c'' and define an "effective" diffusion coefficient $\mathfrak{D}'$ by

$$-\mathfrak{D}'(c' - c'')/l = \mathfrak{D}\partial c/\partial x \tag{246}$$

If this equation be multiplied by dx and integrated from $x = 0$ to $x = l$

$$\mathfrak{D}'/\mathfrak{D}_0 = 1 + \frac{1}{\Delta c}\int_{c''}^{c'} f(c)\,\mathrm{d}c \tag{247}$$

where $\Delta c = (c' - c'')$. Since $\mathfrak{D}'$ changes as the diffusion proceeds, equation 243 becomes

$$\frac{\mathrm{d}(\Delta c)}{\Delta c} + \beta\mathfrak{D}'\,\mathrm{d}t = 0 \tag{248}$$

and if $F(c', c'')$ be defined by

$$\mathfrak{D}_0/\mathfrak{D}' = 1 + F(c', c'') \tag{249}$$

its integration can be effected by means of the equation

$$\ln\frac{\Delta c_f}{\Delta c_0} + \int_{\Delta c_0}^{\Delta c_f}\frac{F(c', c'')}{\Delta c}\,\mathrm{d}(\Delta c) = -\beta\mathfrak{D}_0 t \tag{250}$$

but since by definition

$$\ln\frac{\Delta c_f}{\Delta c_0} \equiv -\beta\mathfrak{D}_i t \tag{251}$$

$$\mathfrak{D}_i/\mathfrak{D}_0 = 1 + \left(\frac{1}{\beta}\mathfrak{D}_0 t\right)\int_{\Delta c_0}^{\Delta c_f}\frac{F(c', c'')}{\Delta c}\,\mathrm{d}(\Delta c) \tag{252}$$

Recalling that c_i is the concentration at which $\mathfrak{D}_i = \mathfrak{D}$, the second term on the right of this equation must equal $f(c_i)$, because of the original assumption expressed by equation 245.

If $\mathfrak{D}$ is known as a function of c, $\mathfrak{D}'$ can be readily calculated by graphical or analytical integration from the known values of c' and c'' by equation 247. Then equation 252 can be used to evaluate $\mathfrak{D}_i$, since for every value of $(c' - c'')$, $F(c', c'')$ can be computed. The more important reverse calculation of $\mathfrak{D}$ from $\mathfrak{D}_i$ requires a short series of approximations. If, as is usually the case, the inner and outer solution volumes are the same, $\mathfrak{D}$ is identified with $\mathfrak{D}_i$ at the mean of the initial concentrations of the inner and outer solution. These values of $\mathfrak{D}$ may be represented by the empirical equation

$$\mathfrak{D}/\mathfrak{D}_0 = 1 - A\sqrt{c} + Bc \tag{253}$$

where A and B are empirical constants. Then by means of equations 244, 247, and 252 a second series of values of $\mathfrak{D}_i$ is computed. The next step is to substitute these values of $\mathfrak{D}_i$ in equation 253 and to calculate the corresponding values of c_i. The second approximation consists in adjusting the constants A and B with these values of c so that they are identified with the observed integral values $\mathfrak{D}_i$. We shall have occasion to reconsider this method of calculation when actual diffusion coefficients by the conductance and diaphragm-cell methods are compared.

D. RESTRICTED DIFFUSION

1. *The layer analysis method*

The method of layer analysis in which a solute diffuses upward in a closed vertical cell and, after a known time, equal portions of the solution at suitable heights are analyzed is in the category of restricted diffusion. The early measurements of Scheffer (86) and Arrhenius (2), the extended results of "Öholm (75, 76), and the precision study of Cohen and Bruins (12) were made by the method of layer analysis. The best technique was effected in cells in which a shearing mechanism was employed, first to form an initial sharp boundary and secondly for removing the layers for analysis (8, 20, 36, 47, 75, 76, 88, 92, 102). The most precise measurements of this kind were made by Cohen and Bruins, who allowed 0.1 N potassium chloride at 20°C. to diffuse from the lower compartment in an accurately machined apparatus divided into four compartments. After 36 hr. the four compartments were separated by a shearing mechanism and the solutions analyzed by a liquid interferometer.

On the assumption that the diffusion coefficient is not a function of the concentration, its determination may be obtained from the solution of the differential equation of Fick's second law for restricted diffusion given by equation 218. If a is the height of each of the four compartments and n_0 the initial concentration of the solute in the lowest compartment, then when $t = 0$,

$$n = n^0 \text{ from } x = 0 \text{ to } x = a;\ n = 0 \text{ for } x = a \text{ to } x = 4a$$

and at all times

$$\frac{\partial n}{\partial x} = 0 \quad \text{for} \quad x = 0 \quad \text{and} \quad x = 4a \tag{254}$$

With these conditions, equation 218 becomes

$$n = \frac{n^0}{4} + \frac{2n^0}{\pi} \sum_{m=1}^{\infty} \frac{1}{m} \sin \frac{m\pi}{4} \cos \frac{m\pi x}{4a} e^{-(m^2\pi^2\mathfrak{D}t)/16a^2} \tag{255}$$

At a time t, the quantity of solute in each of the four equal layers may be obtained by the integrals

$$Q_p = A \int_{(p-1)a}^{pa} n \, dx \tag{256}$$

where A is the cross-sectional area of the cell.

This method has been thoroughly investigated by Stefan (91) and Kawalki (51), who constructed tables which were regarded by the earlier investigators as suitable for the computation of $\mathfrak{D}$. These give the percentage of the total quantity of solute in each of the layers as a function of $\mathfrak{D}$, t, and a: namely, $\mathfrak{D}t/(\Delta a/2)^2$. Thus, by matching the experimental results with those in the tables, a value of $\mathfrak{D}$ may be obtained.

As shown by Cohen and Bruins (12), this experiment may be carried out with great precision but unfortunately, the values of $\mathfrak{D}$ are not differential diffusion coefficients nor are they the same as integral diffusion coefficients corresponding to a concentration of one-fourth the initial concentration. Further, Gordon (31, 32) showed that the $\mathfrak{D}$ of Cohen and Bruins was not the same as the integral coefficient obtained from diaphragm-cell measurements. Hartley and Runnicles (43) examined the matter in greater detail and came to the conclusion that the coefficient of Cohen and Bruins, obtained by diffusing 0.1 N potassium chloride into three-fourths the volume of water, was approximately the integral coefficient which would be obtained in a diaphragm cell if 0.06 N potassium chloride were allowed to diffuse into pure water. Both Gordon, and Hartley and Runnicles, are of the opinion that the result of Cohen and Bruins is high, owing to mixing when the apparatus is separated into layers.

2. *A conductance method for the differential diffusion coefficient*

The possibility of utilizing conductance measurements for determining the diffusion coefficient of electrolytes has naturally occurred to many investigators, but only very recently has it been carefully developed. Niemöller (73), over fifty years ago, measured the change in conductance through a capillary tube containing a solution of a diffusing electrolyte. Haskell (44) employed a diffusing column 50 cm. in length and 5 cm. wide and recorded the conductance at various heights. Lamm (57) has recently developed a method and described a cell for the determination of the differential diffusion coefficient by this method. Lamm develops the theory of an apparatus in which free diffusion occurs. This method is more complicated than the method based upon restricted diffusion employed by Harned and French (39). The latter demonstrated that their apparatus was capable of yielding differential diffusion coefficients of potassium chloride at concentrations between 0.002 and 0.005 N with an accuracy of the order of 0.9 per cent. As a result of their experience, they stated that with a new design of cell and with some technical improvements this accuracy could be greatly increased. This has proved to be the case (40) and there seems to be little doubt that this method is well calculated to solve experimentally the fifty-year-old problem of diffusion coefficients of electrolytes in very dilute solutions.

The most important feature of the method consists in designing a cell in such a way that mathematical complications are reduced to a minimum and the interpretation of the measurements is simple and direct. The simplest possible form of cell is a rectangular parallelepiped of height a, completely filled with liquid in which an electrolyte is allowed to diffuse upward. This process belongs to the category of restricted diffusion. The calculations are greatly simplified by measuring the difference of conductance between two electrodes near the bottom and two electrodes near the top of the cell, their exact position to be determined by theory. For this idea, we are indebted to Professor Lars Onsager.

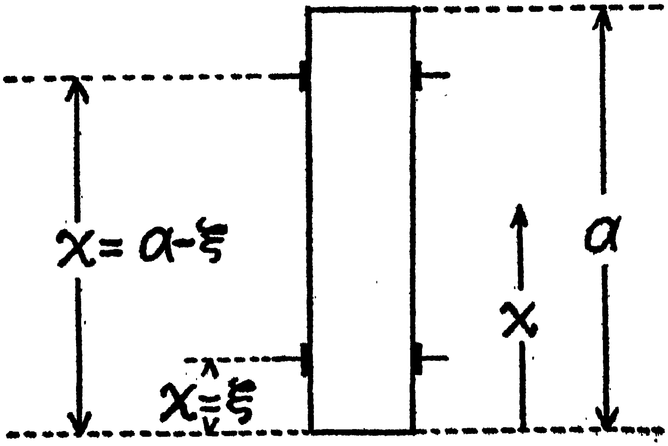

FIG. 8. Vertical cross-section of conductance cell showing quantities involved in the theoretical derivations.

The schematic cross section of such a cell is shown in figure 8, in which the electrodes are at a distance, ξ, from the top and bottom. Solute in the solution from the bottom part of the cell diffuses upwards into the top part, so that as time elapses the concentrations at the top and bottom of the cell approach each other. We shall assume that this difference in concentration is small enough so that $\mathfrak{D}$ may be regarded as constant. Then, according to Fick's second law

$$\frac{\partial c}{\partial t} = \mathfrak{D}\frac{\partial^2 c}{\partial x^2}; \qquad \frac{\partial c}{\partial x} = 0 \quad \text{for} \quad \begin{cases} x = a \\ x = 0 \end{cases} \tag{275}$$

where c is the concentration in mols per liter. These conditions are satisfied by the solution (equation 218) in the form

$$c = \sum_{m=1}^{\infty} A_m e^{-m^2\pi^2\mathfrak{D}t/a^2} \cos \frac{m\pi x}{a} + c_0 \tag{258}$$

where the Fourier coefficients are such as to satisfy the initial conditions of concentration. The difference in concentration at the bottom $c(\xi)$ and the top $c(a - \xi)$ is

$$c(\xi) - c(a - \xi) = 2A_1 e^{-\pi^2\mathfrak{D}t/a^2} \cos \frac{\pi\xi}{a} + 2A_3 e^{-9\pi^2\mathfrak{D}t/a^2} \cos \frac{3\pi\xi}{a} + 2A_5 e^{-25\pi^2\mathfrak{D}t/a^2} \cos \frac{5\pi\xi}{a} + \cdots \tag{259}$$

We note immediately the great advantage in measuring the difference in concentrations at the bottom and top electrodes, since all even terms of the series vanish and, owing to the character of the experimental term, the series converges very rapidly. Now, if the electrodes are placed at a distance ξ from the top and bottom equal to one-sixth the total depth of the cell ($\xi = a/6$), then

$$c(\xi) - c(a - \xi) = 2A_1 e^{-\pi^2\mathfrak{D}t/a^2} \cos \frac{\pi\xi}{a} + 2A_5 e^{-25\pi^2\mathfrak{D}t/a^2} \cos \frac{5\pi\xi}{a} + \cdots \tag{260}$$

and only the first term of the series has significance after sufficient time has elapsed. Since ξ and a are constant,

$$\ln\,(c(\xi) - c(a - \xi)) = -\frac{\pi^2\mathfrak{D}t}{a^2} + \text{constant} = -\frac{t}{\tau} + \text{constant} \tag{261}$$

where $\tau = a^2/\pi^2\mathfrak{D}$, $1/\tau$ becomes the slope of the plot of

$$\ln\,(c(\xi) - c(a - \xi)) \textit{ versus } t$$

and

$$\mathfrak{D} = \frac{a^2}{\pi^2}\frac{1}{\tau} \tag{262}$$

As a result of this design of cell, the equation of the diffusion process becomes one of the first order, so that if the concentration difference at the bottom and top be measured accurately, $\mathfrak{D}$ may be evaluated by the simple measurement of the depth of the cell.

For many electrolytes over the narrow ranges of concentration involved in this experiment the difference in concentrations, $c(\xi) - c(a - \xi)$, is proportional to the difference in conductance, $K_B - K_T$, so that

$$\ln\,[K_B - K_T] = -t/\tau + \text{constant} \tag{263}$$

When this is the case, the calculation of $\mathfrak{D}$ is considerably simplified.

It is apparent that this method is most direct and has the greatest theoretical simplicity. Only the measurements of the resistances at the top and bottom of the cell at suitable times and the depth of the cell are required. It has the advantage of great accuracy in dilute solutions. Its only disadvantage is the length of time required to obtain suitable results.

The procedure of measuring concentrations at the top and bottom of an apparatus during restricted diffusion might be made the basis of an optical method. Two parallel beams of light could be passed through the cell at suitable positions at the top and bottom. By a proper arrangement of an optical sytem, the interference phenomena produced could be used to measure the difference in concentrations at the bottom and top as the diffusion proceeds. Such a method might prove to possess some advantages over other optical methods.

VI. EXPERIMENTAL RESULTS

1. The differential diffusion coefficient of potassium chloride in dilute aqueous solution

The conductance method described in the last section has provided the most accurate values of diffusion coefficients in dilute solutions. The second column of table 2 contains values of $\mathfrak{D} \times 10^5$ of potassium chloride in water recently determined by Harned and Nuttall (40). These results were obtained by an apparatus which differed somewhat from the one used by Harned and French (39) in their earlier measurements by this conductance method. This and other improvements in technique have reduced the error of the determination from ± 0.9 per cent to the order of ± 0.1 per cent.

Each of these determinations depends on three measurements,—differences in conductance between the bottom and top of the cell taken at suitable intervals over a period of 6 days, the times when these conductances were recorded, and the depth of the cell. Values of the diffusion coefficients were calculated from the first to second, second to third, etc. days from the slopes derived from 24-hr. differences in conductance. These values were constant to within narrow limits and, as shown by Harned and Nuttall, showed no trend over a period of 6 days. The results in the table represent the mean of all these individual determinations.

The fifth column contains values of $\mathfrak{D}$ computed by equation 176. All the factors involved are considered. In this concentration range, the calculation $\mathfrak{M}/c$ for potassium chloride is comparatively simple. Since the equivalent conductances of the ions involved are so nearly equal, the second term on the right of equation 178 is negligible. The calculation of the activity term is also simple at low concentrations, since the term $\psi(d)$ in equation 184 may be neglected. Upon introducing the values: $\lambda_1^0 = 73.48$, $\lambda_2^0 = 76.34$ from the data of Shedlovsky (90), and Longsworth and MacInnes (65), $\eta_0 = 8.949 \times 10^{-3}$ and $D_0 = 78.54$, equation 179 becomes

$$\frac{\mathfrak{M}}{c} \times 10^{20} = 40.254 + 18.96c\phi(A'\sqrt{c}) \tag{264}$$

from which the values of $\mathfrak{M}/c$ in column three were obtained. The activity coefficient term in column four was computed by equation 183 upon substitution of the parameters given by Harned and Owen (41): namely, $\mathcal{S}_{(f)} = 0.5065$' $A' = 1.249$, and $B = 0.0202$. The term containing D' is negligible in solutions as dilute as those under consideration.

The agreement of the observed values with the theoretical ones is remarkable. This is illustrated most clearly by the results in the last column of the tables, in which the calculated values of $\mathfrak{D}_0 - \mathfrak{D}$ have been added to the experimental results. If the theory is valid, the resulting values of $\mathfrak{D}_0$ should be constant and equal the limiting value computed from conductance data. The mean value of these results is 1.997, a figure which confirms the limiting value of 1.996. In fact, as pointed out by Gordon (32), these are the first objective results of sufficient accuracy to confirm the limiting equation of Nernst for the differential diffusion coefficient at infinite dilution.

TABLE 2

The observed and calculated values of the differential diffusion coefficient of potassium chloride at 25° C.

c	$\mathfrak{D} \times 10^5$ (OBSERVED)	$\frac{\overline{\mathfrak{M}}}{c} \times 10^{20}$	$1 + c\frac{\partial \ln y_\pm}{\partial c}$	$\mathfrak{D} \times 10^5$ (THEORETICAL)	$\mathfrak{D}_0 \times 10^5$
0.00000		40.253	1.000	1.9958	
0.00125	1.961	40.301	0.981	1.9605	1.996
0.00194	1.954	40.317	0.977	1.953	1.997
0.00325	1.943	40.346	0.971	1.943	1.996
0.00585	1.931	40.395	0.963	1.929	1.997
0.00704	1.924	40.413	0.960	1.925	1.995
0.00980	1.918	40.449	0.955	1.916	1.999

The magnitudes of the electrophoretic contribution and the activity coefficient terms are illustrated by figure 9. The lowest plot represents the limiting law as calculated by equations 185, 186, and 187, which upon substitution of the experimental values indicated reduces to

$$\mathfrak{D} \times 10^5 = 1.9958 - 1.170\sqrt{c} \tag{265}$$

The dotted curve was computed by equation 177 by employing the limiting value of $\overline{\mathfrak{M}}/c$ and neglecting its change with concentration. This leads to the numerical equation

$$\mathfrak{D} \times 10^5 = 1.9958 \left(1 + c\frac{\partial \ln y_\pm}{\partial c}\right) \tag{266}$$

The graph at the top is the plot of the results of the complete theoretical computation given in the fifth column of table 2. The circles are the observed values from the second column of this table.

There is little doubt from the agreement between the theoretical and observed results that, in these dilute solutions, the term in equation 264 containing $\phi(A'\sqrt{c})$ is required. Indeed, this is the first experimental confirmation of the Onsager and Fuoss theory.

2. Comparison with diaphragm-cell measurements[12]

The only other measurements at low concentrations which can be compared with these results have been obtained by the diaphragm-cell method. In his recent contribution, Gordon (32) gives the results of his computation of the differential diffusion from the integral values by the method outlined in Section V, C. The experimental data of Hartley and Runnicles (43), McBain and Dawson (67), and Gordon were used. The selection of a standard for the calibration of the diaphragm cells depended on the semi-empirical equation of Gordon, which represented Clack's data at the higher concentrations and which extrapolated to the Nernst limiting value. The results at low concentrations fall closely to the dotted line in figure 9.

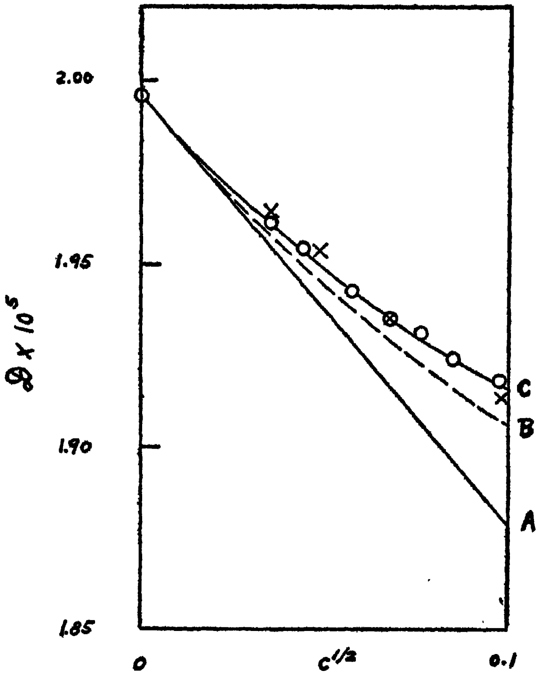

FIG. 9. The diffusion coefficient of potassium chloride in dilute aqueous solution at 25°C. O, conductance method; ×, diaphragm-cell method. A, limiting equation 265; B, equation 266; C, complete theoretical computation.

These results have now been revised by Gordon upon the basis of the new measurements given in table 2. This is much more satisfactory, since the calibration depends on objective evidence only. The data in this table are represented accurately by the empirical equation 253, which with numerical values becomes

$$\mathfrak{D} \times 10^5 = 1.996 - 1.065\sqrt{c} + 2.75c \qquad (267)$$

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[12] The subject matter and calculations of this section have been derived from a private communication from Professor A. R. Gordon of the University of Toronto.

On this basis and the method of computation described in Section V,C, the results in table 3 were obtained. The values of $\mathfrak{D}_i$ are differential diffusion coefficients at the concentration c_i, since c_i is the concentration at which $\mathfrak{D}$ and $\mathfrak{D}_i$ are equal. These results are plotted as crosses in figure 9. The value at 0.00468 represented by the circle and cross was used for calibration. The other results agree within 0.004 with those obtained by the conductance method. This agreement is probably within the experimental error, and nothing further need be said until further data are available in dilute and moderately dilute solution.

TABLE 3

The diffusion coefficient from diaphragm-cell measurements

$c_0'' = 0$; $V' = V''$; $\Delta c_f/\Delta c_0 = 0.5$

c_0'	$10^5\mathfrak{D}_i$ (OBSERVED)	c_i
0.10	1.838	0.044
0.05	1.871	0.0225
0.025	1.904	0.0115
0.02	1.914	0.0093
0.01	(1.936)*	0.00468
0.005	1.954	0.00235†
0.0025	1.968	0.00118†

* Calibration of the diaphragm cell was based on this concentration.

† From the data of Hartley and Runnicles.

3. *Comparison of experimental and theoretical results at high concentrations*

The characteristics of the diffusion coefficient in concentrated solutions in relation to the theory are illustrated in figure 10, where $\mathfrak{D} \times 10^5$ for potassium chloride in water at 25°C. is plotted against $c^{1/2}$. Here the curves shown in figure 9 are extended to high concentrations. Curve A represents values computed by the limiting equation (265), curve B those given by equation 266, and curve C those obtained by a complete calculation of $(\overline{\mathfrak{M}}/c)$. In concentrated solutions, the term containing the function $\psi(d)$ in equation 183 contributes as large an amount as the $(\overline{\mathfrak{M}}/c)$ term.[13] The circles represent the conductance results shown in figure 9 and the crosses results computed by Gordon (32).

The existence of the minimum in this plot which corresponds closely to the minimum in the activity-coefficient function confirms the premise that the gradient of the thermodynamic potential is one of the important factors in concentrated solutions. This is confirmed by the data for other electrolytes. However, in the case of potassium chloride as well as all other electrolytes so far examined, the experimental results above 0.01 N lie considerably below those resulting from theory. This is not surprising for, as pointed out in Section IV,6, the obscure ion–solvent interactions and any other short-range repulsive force interactions between the ions are not yet subject to quantitative treatment and were omitted in the development of the theory.

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[13] This is also true for sodium chloride in aqueous solution at 18.5°C., as shown by Harned and Owen (41).

Onsager and Fuoss suggested that the change in viscosity caused by the electrolyte would have an appreciable effect on the $\overline{\mathfrak{M}}/c$ term in the more concentrated solutions and that as a first approximation should be inversely proportional to the macroscopic viscosity. Harned and Owen (41) have shown that for sodium chloride solutions at 18.5°C. multiplication of $\overline{\mathfrak{M}}/c$ by η_0/η causes the theoretical curve to approach the experimental at moderate concentration but overcorrects in solutions of higher concentrations. Deviations from the theory and the viscosity factor have been examined in considerable detail by Van Rysselberghe (100).

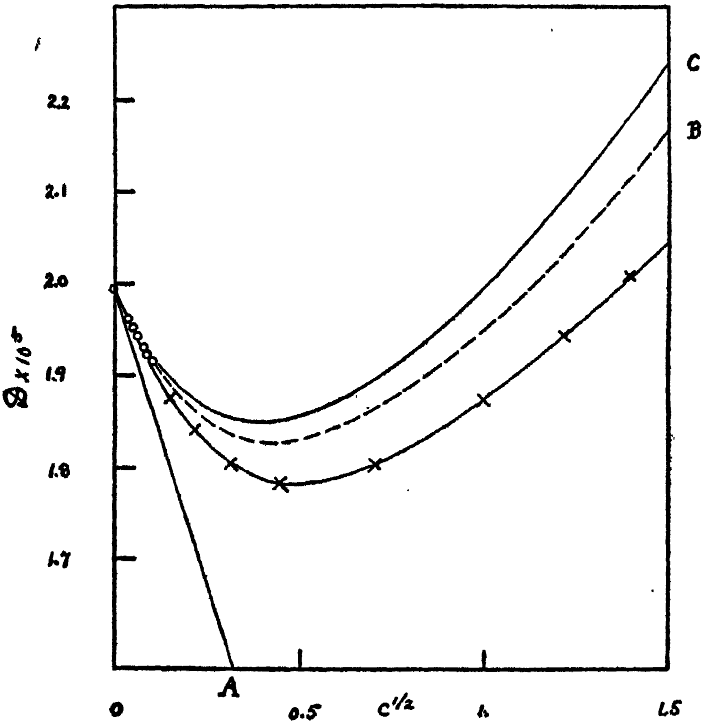

FIG. 10. The diffusion coefficient of potassium chloride at 25°C. O, conductance; × diaphragm cell. A, limiting equation 265; B, equation 266; C, complete theoretical computation.

The empirical equation of Gordon (31) is

$$\mathfrak{D} = 2000RT\left[1.074\times10^{-20}\,\frac{\overset{0}{\lambda}_+\overset{0}{\lambda}_-}{\Lambda^0}\,\frac{V}{n_1\bar{V}_1}\,\frac{\eta_0}{\eta}\right]\left(1+m\,\frac{\partial\ln\gamma_\pm}{\partial m}\right) \qquad (268)$$

in which both the viscosity and the partial molal volume of the solvent are introduced in the mobility term. This reduces to the Nernst equation at the limit but leads to a result in dilute solutions lower than that obtained by the theoretical equation (equation 176). For sodium and potassium chlorides and for potassium nitrate in water, this equation fits the results of Clack (9, 10) from 0.05 to 2 N with considerable accuracy. This is interesting, since at 18°C. η_0/η is less than unity for sodium chloride and greater than unity for potassium chloride and nitrate solutions. For higher valence type electrolytes it does not represent the results as accurately.

VII. GENERAL CONSIDERATIONS, CRITICAL OBSERVATIONS, AND SUMMARY

The preceding discussion has incorporated the fundamental theoretical derivations upon which criticism and further developments may be expected. The theory of reciprocal relations in irreversible processes has provided certain symmetry conditions which are fundamental to the specialized theories of conductance of electricity, heat, or matter in solutions. In his most recent paper, Onsager (79) has made a notable contribution to the theory of diffusion in systems of more than two components. For a system of this kind Fick's law may be generalized to

$$\mathrm{J}_i = -\Sigma_k D_{ik} \nabla c_k \tag{269}$$

which defines a set of diffusion coefficients. Interrelations between members of this set are deduced, the equations for the dissipation function are stated, and relations between the diffusion coefficients and thermodynamic properties for a system of more than two components are discussed. For diffusion in electrolytic solutions of three or more components, Onsager (79) has introduced some new substitutions and expressed the part of the dissipation function due to the relaxation effect in a form which is simpler than that formerly obtained by Onsager and Fuoss (80), which required the solution of secular equations. Since there are no experimental data for multiple-component diffusion, these considerations go beyond the scope of this review.

Our descriptions of the methods of measurement have been limited to those which at present seem to be most promising. So far we have found that the Nernst limiting law has been verified by only one series of results on potassium chloride in water obtained by the conductance method. These results are also the only ones which indicate that part of the mobility term resulting from electrophoresis is required in interpreting the data in dilute solutions. We note that in this calculation the term in equation 179 which contains $(\lambda_1 - \lambda_2)^2$ is negligible, and consequently no evidence from any source is available to prove that this term is significant.

The only other results of comparable accuracy with these conductance measurements have been derived from diaphragm cells which have the disadvantage of being relative. It is interesting that for hydrochloric acid the limiting value of the diffusion coefficient as estimated by James and Gordon (49) from diaphragm-cell measurements differs by 6 per cent from the value computed by the Nernst equation.

Departures from the theoretical equation for potassium chloride occur in the neighborhood of 0.01 N and wide departures are found in concentrated solutions —where terms in powers of the concentration higher than $c^{1/2}$ become important. All measurements in both dilute and concentrated solutions confirm the importance of the activity coefficient term in the theoretical equations. In view of our lack of knowledge of concentrated solutions (e.g., ion–solvent etc. interactions), these deviations from the theory are inevitable. Indeed, much more data on electrolytes will be required to prove or disprove the general validity of the theory at very low concentrations. Further advances will depend on whether a quantitative statistical theory of such interactions can be developed.

Hermans (46) criticizes the theory of Onsager and Fuoss (80) on the grounds that it is illogical to employ a term of higher power than unity in the Boltzmann equation (128), while neglecting such a term in the Debye and Hückel theory (127). He allows no terms greater than a first power of $c^{1/2}$. This leads to the elimination of the term containing $\phi(A'\sqrt{c})$ in equation 179. In media of high dielectric constant, however, the effect of dropping the higher terms of the Debye and Hückel theory is small as compared to the $\phi(A'\sqrt{c})$ term, so that the Onsager and Fuoss procedure is justified. The diffusion coefficient of potassium chloride indicates definitely that this term is required, but it would be a mistake to rely solely on one series of measurements as final proof of a matter of this kind.

In addition to the data of Clack (9, 10) and Davies (15) in concentrated solutions, mention should be made of the work of James and Gordon (49) and Hollingshead and Gordon (48), who determined the differential diffusion coefficients of hydrochloric and sulfuric acids as a function of temperature. They find that the activation energy is a function of the concentration for hydrochloric acid above 0.2M but is not a function of the concentration for sulfuric acid from 0 to 1 M.

This review began with the statement that there are few fields of physical science which, for the effort made, have yielded so few experimental results of accuracy sufficient to test any part of the theory. This includes the limiting equation of Nernst, the validity of which has never been questioned but which with the exception of one series of results shown in figure 9 has never received experimental confirmation. By the use of the conductance method and the diaphragm-cell method in dilute solutions, it seems probable that many of the problems in very dilute solution may be solved. In more concentrated solutions optical methods may be easier and preferable. At present, there seems to be little chance of the development of a quantitative theory of concentrated solutions, but perhaps the nature of the deviations from theory in dilute solutions may prove a help to future advance.

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